LABORATORY MANUAL OF CHEMISTRY

BRADBURY

ELEMENTARY CHEMISTRY

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PART II
EXPERIMENTAL WORK

BY

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PREFACE

In the preparation of the Laboratory Manual I have tried to make the directions as complete and clear as possible. Wherever there is any danger in carrying out an experiment, the necessary precautions are explicitly stated, even at the cost of some repetition.

Most of the experiments—but not all of them—are intended for individual laboratory work. This is a matter which will vary with the equipment of the school and the size of the classes. The following is a list of the numbers of the experiments which I have found to be better fitted for lecture-table or laboratory demonstrations than for individual work, but it is merely the result of my own experience, and is not intended to have any prescriptive implication:

Numbers 7, 9, 10, 27, 31, 40, 49a, 50, 51, 60, 128, 129, 133.

The questions placed after each chapter are based partly on the laboratory exercise and partly on the corresponding chapter of Part I. The student should not fall into the error of supposing that his own experimental work supplies all the necessary data for answering them.

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* Chapters correspond to those in Part I.

ELEMENTARY CHEMISTRY

PART II

EXPERIMENTAL WORK

General suggestions.—Chemical laboratory work is by no means free from danger. The eyes, especially, are likely to be injured if the work is not done with proper care. The secret of safety is the accurate observance of the directions given for the performance of the experiment. Before carrying out an experiment the student should carefully read the directions and should observe them scrupulously, asking for information on any points which are not clear.

Neatness is essential to success. Experiments made in dirty test-tubes, beakers, or mortars, are worthless, misleading, and often dangerous. Every piece of apparatus employed must be spotlessly clean and must be cleaned carefully after the work is done.

The sinks in the laboratory tables are intended only for liquids. Solids, like bits of broken glass, match-sticks, and paper, must never be thrown into them.

The student must be continually on his guard against the tendency—almost universal with beginners—to take *too much* of the various chemicals required in the experiments. One objection to this is waste of material. Another is waste of time, for it always requires a much longer time to carry out a process with a large quantity of substance than with a small quantity. Finally, there are many

experiments which are perfectly safe on the small scale but become highly dangerous when the quantity is increased. As a rule, the quantity of material to be taken is indicated in the directions, and nothing but loss of time and danger result from taking more. Where no amount is stated, take the smallest that you can conveniently work with.

Anything which is spattered into the eyes must be removed instantly by copious washing with water. Acid splashed upon the skin should be washed off at once with abundant water, and if a wound has resulted from its action, it should be treated with a paste made of sodium acid carbonate (baking soda) and water. The same paste may be applied to burns.

Preliminary Exercise.—The Bunsen Burner; Glass Rod and Tubing

A. **The Bunsen burner.**—Examine the burner. Close the holes at the base and light it. Hold a piece of glass tubing in the flame for a time. Open the holes and describe the change which the flame undergoes. Hold a piece of glass tubing in this flame. Which flame is hotter? Which is cleaner? Which is least affected by drafts? What is the cause of the difference between the two flames? Which is the hottest part of the blue flame?

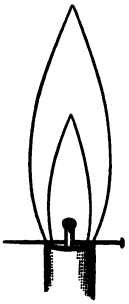

Fig. 1.

Is the blue flame hollow or solid? Obtain facts to answer this question by thrusting a match-stick horizontally through the flame near the burner. Support a match by means of a pin, as shown in Fig. 1, and light the burner. Remove the match, and hold in the center of the blue flame near the burner one end of a glass tube open at both ends. The tube should be inclined obliquely upward. Hold a lighted

match to the upper end of the tube. What does this show regarding the nature of the interior of the flame?

Open the holes at the base as wide as possible, and gradually turn off the gas until the flame strikes back. Why does the flame only strike back when the current of gas up the chimney becomes slow? While the flame is still burning below, turn on the gas and light it above. Is this flame suitable for use? How can the burner be made to give the blue flame again?

The flame must never be allowed to burn below, since it gives off poisonous gases. Remember in using the burner that—except where high temperatures are required —a small flame is better than a large one. Do not light the burner until you are ready to use it, and always turn it down or extinguish it when it is not in use.

Take the burner apart, make drawings of the parts, and explain the function of each.

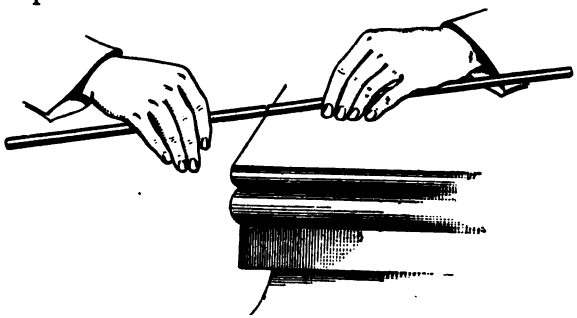

FIG. 2.

B. **Glass rod.**—With a triangular file make a notch on a piece of glass rod 15 centimeters[1] (6 inches) from the end. One sharp stroke of the file is sufficient. Holding the rod as indicated in Fig. 2 endeavor to bend it away from the notch, and it will break off at that point.

[1] Hereafter the abbreviation cm. for centimeters will be used.

Cut off three such pieces. Since the ends of the rods are jagged and inconvenient to handle, round both ends of each rod by holding it in the Bunsen flame and rotating the rod gently. Support the rods by the middle on your test-tube rack until they cool. (Hot glassware or hot apparatus of any kind must never be laid on the desk or put away under it.)

C. **Glass tubing.**—Examine a piece of hard and a piece of soft glass tubing. Carefully note and record the differences in appearance between them. Select for yourself a piece of each kind of glass from the main stock. Cut, just as you cut the glass rod, two pieces of the soft glass, each 20 cm. (8 inches) in length. Cut one piece of hard glass tubing of the same length.

Hold the middle of your piece of hard glass tubing in the flame, turning it gently, and when it becomes red-hot, gently and slowly draw the two portions asunder. Do not twist the tubes. The pull must be straight. Let the two tubes cool and use them to study the effect of heat upon a fragment of wood (match-stick) and a little sugar.

In the same way make four sealed tubes of soft glass. The temperature required is not so high, and the glass must be removed from the flame before drawing it out. Then the thin middle portion must be returned to the flame and melted and the two tubes formed drawn apart.

Use these tubes for studying the effect of heat upon a little paper, a crystal of iodine, a fragment of sulphur, and a small piece of starch. *Never heat anything in a tube sealed at* BOTH *ends, since this would cause explosions.* Soft glass tubing is used for all ordinary purposes; and hard glass tubing—which is much more expensive and more difficult to manipulate—only when high temperatures are to be applied.

Bending glass tubing.—For bending use the flame of a wing-top burner, never the Bunsen flame. Hold the tube—which should be of soft glass—so that the flame heats as long a portion as possible, and rotate it so that it is evenly heated. When sufficiently hot remove it from the flame and make the bend. When the tube is perfectly cold remove the soot by wiping the outside with paper.

In this way, bend a soft glass tube into an acute angle (Fig. 3). Select a piece of soft glass tubing about 50 cm. (20 inches) long and bend it twice at right angles

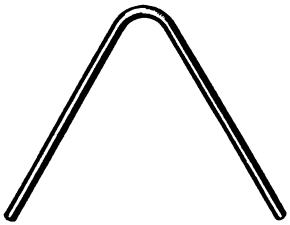

FIG. 3.

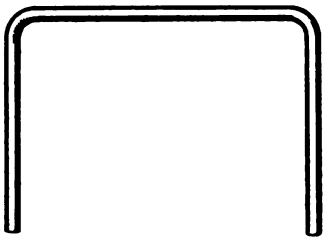

FIG. 4.

(Fig. 4). This must be done so that when the double bend is laid upon the table every part of it will touch the surface of the latter, or, in other words, the two limbs must be in the same plane. Round the sharp ends of the bent tube by holding them a short time in the flame—not long enough to cause them to collapse.

CHAPTER I

WATER

EXPERIMENT 1.—Place a small piece of potato in a dry test-tube. Clamp the tube horizontally and heat gently with a small flame. As soon as a positive result is obtained stop the experiment.

What is the result? Other vegetable substances would behave in the same way. Draw a conclusion regarding the existence of water in living things. Mention some vegetable products in which the presence of abundant water is evident on mere inspection.

EXPERIMENT 2.—Crush some ice in a mortar and half fill a small round-bottomed flask with it. Dry the outside of the flask carefully with a towel and clamp it about 25 cm. (10 inches) above the desk. Directly under it place a clean, dry beaker. Let the apparatus stand for an hour or more while you go on with other work. What does the result prove? How does warm air differ from cold air in its capacity for water? What is dew and how is it formed? What is humidity and how does it affect our sensations in warm weather? Suppose the flask had contained a mixture of ice and salt, what difference would have been noticed in the result? If water should suddenly cease to evaporate into the air, what changes would occur upon the earth's surface? Sketch the apparatus in your note-book.

EXPERIMENT 3.—Carefully evaporate a little faucet water to dryness in a perfectly clean porcelain dish. Is the water pure? If not, what is the source of the impurities? Evaporate a little distilled water in the same way. What is the difference?

EXPERIMENT 4.—Set up the apparatus shown in Fig. 5 and distill some water in it. The flash should be half filled with water which has been colored by ink. It must be perfectly dry on the outside, and is heated gently with a piece of wire gauze between it and the flame to avoid breakage. Before heating, connect the condenser with the water supply and pass a gentle current of water in at the lower tube, letting it run off at the upper tube into the sink. Avoid violent boiling, which would carry over the impurities into the receiver. Distill about 100

cubic centimeters [1] of water, stopping before the water in the flask becomes very low, otherwise the latter would break. Test the purity of the water by evaporating a small quantity in a perfectly clean dish.

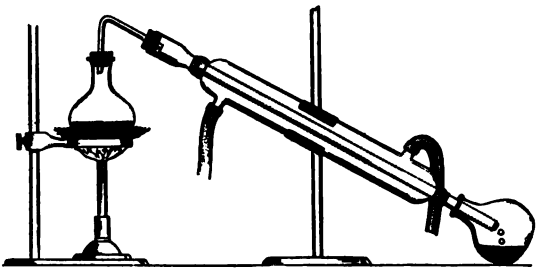

FIG. 5.

How does distillation remove the impurities from water? Can *all* impurities be removed in this way? Why is glass apparatus not used in making distilled water commercially? Make a drawing of the apparatus.

EXPERIMENT 5.—Fill a beaker with crushed ice and clamp a thermometer so that the bulb is surrounded by the mass. Notice the behavior of the mercury column and, when it becomes stationary, take several readings. Take the average of the temperatures and record it as the result of a determination of the melting-point of ice.

Suppose you had placed liquid water in the beaker; had surrounded the beaker with a freezing mixture and taken the temperature when the water was partially frozen, would the result have been the same?

What evidence have we that ice and water are two forms of the same substance?

Make a mixture of salt and crushed ice and take the temperature as before. (?)

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[1] c.c. will be used hereafter as the abbreviation for cubic centimeters. 100 c.c. is about one-half the capacity of a kitchen cup.

EXPERIMENT 6.—Half fill a small beaker with distilled water, dry it on the outside and support it on wire gauze. Clamp a thermometer so that the bulb is in the liquid and heat. Watch the thermometer. Small bubbles escape before the liquid begins to boil. Explain this. What would be the effect of putting fish in water which had been boiled and then cooled in the absence of air?

When the liquid boils, take several readings of the temperature. Record the results. What is the effect of variation of pressure on the boiling-point of water. Why is it difficult to cook certain kinds of food on high mountain-tops?

With a pipette measure off 50 c.c. of water into a 100 c.c. beaker. Dissolve in the water 10 grams of salt and take the boiling-point of the solution. Repeat, using 10 grams of sugar. How do the boiling-points of solutions of solids compare with that of pure water? Does sugar or salt more strongly affect the boiling-point?

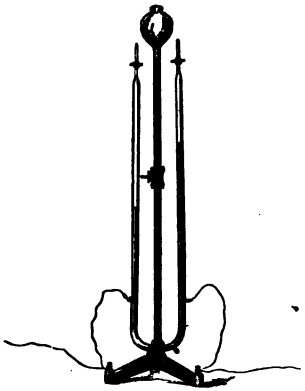

FIG. 6.

EXPERIMENT 7. — Carry out the *electrolysis of water* in the apparatus shown in Fig. 6. Dilute some sulphuric acid with about ten times its volume of water. *In diluting sulphuric acid pour the acid into the water in a thin stream, stirring constantly—never the water into the acid.* (With the other acids it is a matter of indifference.) Cool this liquid by standing it in a tin pan full of water. Fill the apparatus with the liquid by pouring it into the funnel tube until it just reaches the stopcocks. The latter must be open during this operation.

Close them and connect the two pieces of platinum foil in the apparatus with a battery or a dynamo-circuit. Allow the current to pass until about 15 c.c. of the gas which is formed in smaller quantity has collected. Do not let the level of the liquid fall as far as the platinum foil.

Stop the current, and if the apparatus is graduated, read off carefully the volumes of both gases. *Read from the bottom of the meniscus* (Fig. 7). If the apparatus is not graduated, measure the length of each gas-column with a meter scale. What is their relation by volume?

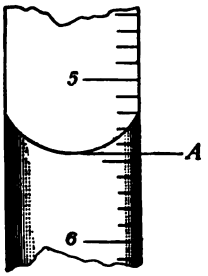

FIG. 7.

Cautiously open each stop-cock for an instant to drive out any water which may be above them. Obtain a pine splint bearing a spark by extinguishing the flame of the burning splint by a quick movement of the hand. Let the gas which is present in smaller quantity stream out against the spark. (?) Slip over the end of the other tube a very short piece of rubber tubing connected with a short piece of glass tubing drawn out to a fine jet. Let the gas escape against the flame of a burning match, and remove the flame. (?) What evidence does this experiment furnish of the composition of water? Why does the level of the liquid in the apparatus sink so slowly when it is run with the stop-cocks open, although much gas escapes? What becomes of the sulphuric acid and what is the object of adding it? (Part I, p. 6).

EXPERIMENT 8.—**Qualitative synthesis of water.**—Fill a U-shaped tube with fragments of dry calcium chloride, free from powder. Cork the open ends of the tube tightly and connect it on the one side with a Kipp generator furnishing a current of hydrogen, and on the other with a

short tube of hard glass drawn out to a jet (Fig. 8). Pass a gentle current of hydrogen through the tube for a few

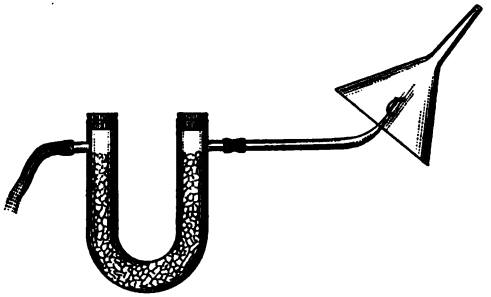

FIG. 8.

seconds to drive out the air, and then light the gas at the jet. Hold over the flame a cold, dry funnel. (?) What is the conclusion? Make a drawing of the apparatus. The object of the calcium chloride is to dry the gas. Why? Hold a cold, dry cylinder or bottle over a burning candle for a moment. Over a small gas flame. (?) What evidence does this furnish with respect to the composition of the gas and the candle?

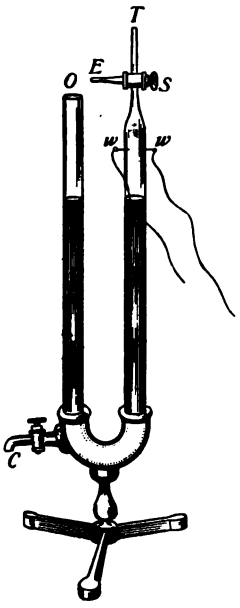

FIG. 9.

EXPERIMENT 9.—**Quantitative synthesis of water.**[1]—This experiment is made in the U-shaped eudiometer (Fig. 9). The end at *O* is open. The other limb of the

[1] This experiment requires some expertness in manipulation. It will be well to carry it out on the lecture-table at this stage, or to postpone it until the student has acquired some skill in the handling of apparatus.

U is closed by a three-way stop-cock *S*. This stop-cock is so constructed that gas passed in through *T* will either escape into the air by the stem *E* or pass down into the eudiometer, according to the position of the stop-cock. Take out the stop-cock and inspect it until you understand why this should be the case. If necessary smear a *little* vaseline thinly over it, put it back in place and turn it until the vaseline is evenly distributed. This will make it air-tight, which is essential to the success of the experiment.

Open *S* and close the stop-cock at *C*. Pour mercury in at *O* until the apparatus is filled to the level of *E*, but not above it.

Fit a test-tube with a perforated cork and delivery tube (Fig. 10). Over the end of the delivery tube slip a short piece of rubber tubing which will readily slip over the stem *T*, making an air-tight junction. Fill this test-tube to the depth of 2 cm. ($\frac{1}{4}$ inch) with a mixture of 1 part manganese dioxide and 3 parts potassium chlorate. Heat very gently, proceeding exactly as directed in Experiment 23. From time to time test the gas given off through the delivery tube with a splint bearing a spark. When the spark bursts into flame, adjust the three-way stop-cock of the eudiometer so that gas admitted at *T* will escape at *E*, and let oxygen stream through the tip and stop-cock for ten or fifteen seconds. (Why?) Then turn the stop-cock so that the oxygen passes into the tube over the mercury and allow 5 to 8 c.c. of oxygen to pass in. Turn the stop-cock so that the oxygen in the eudiometer is shut off both from the external air and from the oxygen-generator, and *at once* remove the latter.

FIG. 10.

Place a dry beaker under the stop-cock *C*, and allow mercury to run out until the level of the mercury in the two limbs is the same.[1] (Why?) Be careful in all these manipulations not to touch with the hands that part of the tube containing the oxygen. (Why?) Read and record the volume of the oxygen. Repeat the reading.

Now transfer your cork and delivery tube to another test-tube in which you have placed a few pieces of granulated zinc, about 3 cm. (1 inch) of water over them and enough sulphuric acid—added very gradually—to produce a rather brisk evolution of hydrogen.[2] Turn the stop-cock so that gas passed in at *T* will escape at *E*. *Be careful in doing this not to throw the oxygen into communication with the air, as this would spoil the experiment.*

Slip the rubber tube over *T* and allow the hydrogen to stream out through *E* for about two minutes. (Why?) Pass about three times as much hydrogen as you have taken of oxygen into the eudiometer, turn the stop-cock so as to shut it off from the air and the hydrogen generator, and remove the latter. Equalize the levels of the mercury in the two limbs, read off the volume of hydrogen added, and let the apparatus stand several minutes to give the two gases time to mix thoroughly.

The two gases can now be caused to combine by a spark passed between the platinum wires *w w*. The spark can be obtained from an induction coil excited by three Edison-Lalande or by two dichromate cells. Since the explosion may possibly be violent enough to wreck the apparatus, it is well to take the precaution of running out mercury at *C* until the level of the mercury in that limb

[1] If by mistake too much mercury is run out, add mercury at *O* until the level is higher in that limb, and again run out through *C* until the levels are the same.

[2] The hydrogen is best taken from a Kipp apparatus if one is at hand.

is much lower than in the other. This dilutes the gases and reduces the violence of the explosion. Before passing the spark, put the thumb tightly over *O* to prevent any mercury being thrown out.

Let the apparatus stand for five minutes, equalize the levels of the two mercury-columns and read off the volume remaining. Record it. Pour mercury into *O* and show that the residual gas is hydrogen by letting it escape at *T* and burning it. Slip a short rubber tube bearing a jet over *T* before doing this, otherwise the heat of the flame would crack the apparatus. Make the calculation as in the following example:

Suppose you have taken..... 7 c.c. oxygen,
and that the volume after adding hydrogen is 27 c.c.;
then the volume of hydrogen added is.... 20 c.c.

Suppose further that the volume after the explosion is 6.1 c.c.

This residual gas is shown by investigation to be hydrogen. It is clear that 20 — 6.1 or 13.9 c.c. of hydrogen must have disappeared with the 7 c.c. of oxygen, to form water. Hence the relation by volume in which the gases combine is 7 : 13.9 or 1 : 1.98.

Make a drawing of the apparatus in your note-book and give a complete description of your work. Explain the calculation. What becomes of the water formed in the explosion and why is its volume not considered? Is the composition of water always the same? If so, how do you account for the fact that different experiments, even when carried out with great care, always yield slightly different results?

The more numerous and careful the experiments the more closely the average result of them will approach the ratio 1 : 2.

What general statement can be made about the volumes in all cases in which two gases combine?

How can you calculate the composition of water by weight from its composition by volume?

EXPERIMENT 10.—**Action of oxygen on copper. Synthesis of water from hydrogen and copper oxide.**—Fit up the apparatus shown in Fig. 11. *T T'* is a piece of hard

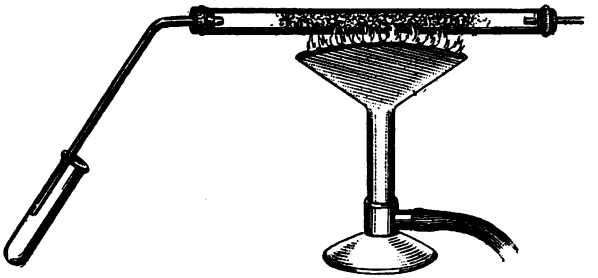

FIG. 11.

glass tubing, 1 cm. or more in diameter and about 30 cm. (1 foot) long. In order to cut off a suitable piece of tubing, first make a notch on the tube at the point where it is desired to cut it. Then wrap several layers of wet filter paper around the tube quite close to the notch, and do this also on the other side of the notch (Fig. 12), so that a narrow band of glass, 1 cm. wide or less, with the notch is left bare. If this bare portion is held in the burner flame, it will usually break off evenly, and any projecting portions can be removed by careful chipping with a key.

FIG. 12.

Thrust a loose plug of asbestos into the tube about 5 cm. (2 inches) from one end, and upon this pour a column of copper clippings about 20 cm. (8 inches) long. Put a similar loose plug of asbestos in the other end of the tube to hold the clippings in place. Holding the tube horizontal, tap it gently upon the table so as to cause the

clippings to settle, and make a channel for the passage of gases.

Select two corks a trifle too large to fit in the ends of the tube, and roll them carefully under the foot until they fit. Perforate the corks either with a rat-tail file or a cork-borer, and fit them with tubes, as in the cut. In perforating with a rat-tail file, thrust the sharp end of the file through the middle of the cork, beginning with the small end, and resting the cork on the table. Then enlarge the hole with the rough part of the file until the glass tube fits it *tightly*. The glass tubes must be rounded at the ends, and must be introduced with a twisting motion and without strong pressure, the use of which is likely to break the tube and seriously cut the hand.

Clamp the tube, or suspend it by a wire, and apply heat from a wing-top burner, very gently at first. If clamps are used, they must be placed at the ends, so as not to be burned. Raise the temperature gradually until the copper is at a dull red-heat. If the tube softens or shows signs of bending, moderate the heat.

Now connect the short straight glass tube with a test-tube in which oxygen is being generated from potassium chlorate and manganese dioxide (Experiment 23), and pass oxygen through it. What is the result? Describe the product. What is its composition? At which end of the tube does the change begin, and why?

Remove the oxygen generator and connect the tube with a hydrogen generator. The Kipp apparatus is best. *If an ordinary generator is used, hydrogen must be allowed to escape briskly from the exit tube for at least five minutes before connecting it to the red-hot tube.* The hydrogen must be passed through a U-shaped tube containing dry calcium chloride in lumps—not powder—to free it from water before it enters the copper tube. See Experiment 22 for the method of generating hydrogen.

What change takes place in the black substance? Hydrogen has a similar effect on the compounds of oxygen with many metals—for instance, upon tin oxide and lead oxide. The effect is called *reduction*. Some other oxides, like those of magnesium and aluminium, are not affected by it.

Look in the cooler part of the tube, and in the bent tube, for evidence of the formation of *water*. What evidence of the composition of water is furnished by this experiment?

Problems

1. The electric current is passed through water containing sulphuric acid until 20 grams of the water have disappeared. What weights of hydrogen and of oxygen have been liberated?

2. 300 grams of water are decomposed. What weights of hydrogen and oxygen are produced?

3. If 27.396 grams of water contain 24.352 grams of oxygen, what is the percentage composition of water?

4. A quantity of hydrogen gas weighing 4 grams is slowly passed through a glass tube containing a large quantity of copper oxide heated to redness. (*a*) How much will this tube lose in weight? (*b*) What weight of water will be produced?

5. 50 c.c. of hydrogen are mixed with 50 c.c. of oxygen and the mixture exploded. (*a*) If the process takes place below 100°, which gas remains in the tube, and how much? (*b*) If the whole process takes place above 100°, what is the total volume remaining in the tube, and of what does it consist?

CHAPTER II

SOLUTION

Experiment 11.—**Soluble and insoluble.**—Place a little coarsely powdered rosin in a test-tube, cover it with water, and shake it. (?) Pour out the water, cover the rosin with alcohol, and shake. (?) Reserve this liquid.

Shake up some copper sulphate, first with alcohol, and then with water. (?) Give definitions of the terms soluble, insoluble, and solution.

Pour the alcoholic solution of rosin into a large quantity of water in a bottle or cylinder. (?) The resulting state of things is called *a suspension.* State some differences between solutions and suspensions.

EXPERIMENT 12.—**Saturated solutions.**—Half fill a liter flask with water and add small quantities of salt, shaking after each addition until the liquid will dissolve no more. Cork the flask and shake the liquid containing a little undissolved salt for a long time, to make sure that it will not take up any more. Pour off the liquid into a narrow cylinder and take its density by means of a *hydrometer* (Fig. 13). Most solutions are denser than water. How about salt-solution? The density will serve as an indication of the amount dissolved. Of course if two solutions of table-salt have the same density, they contain the same percentage of dissolved salt.

FIG. 13.

Half fill the flask again with water and add a large quantity of salt, several times as much as the water can dissolve. Shake for some time. The liquid should contain a sediment of undissolved salt more than 5 cm. in depth. Finally, let the liquid settle, pour off into a narrow cylinder, and take the density as before. The result should be the same.

This important experiment is an example of the fact that *the quantity of a substance which dissolves to produce a saturated solution does not depend upon the excess of undissolved solid present in the liquid.* There must be some solid present; otherwise the liquid would not be saturated; but a solution which remains in contact with a small quantity of a solid without dissolving any more of

it, is just as strong as one which is in contact with a large quantity.

EXPERIMENT 13.—**Effect of temperature on solubility. Crystallization.**—Powder some potassium chlorate in a mortar, not by pounding it, but by moving the pestle with a strong, steady pressure. Place 50 c.c. of water in a 100 c.c. beaker and add potassium chlorate to it in small quantities. Stir until each portion is dissolved before adding the next.

When the liquid is saturated, place the beaker on wire-gauze on a ring of your stand and heat it gently. Continue stirring and adding more potassium chlorate. What is the effect of temperature on solubility? Stand the beaker aside to cool. (?) What is the most obvious difference between a crystal and a bit of non-crystalline matter?

This is by no means the only difference. The great distinction is that all properties which *have direction at all,* are different in different directions in a crystal. Thus, a crystal breaks more easily in some directions than in others ; it conducts heat better in certain directions ; it transmits light quite differently along certain lines, and so on. Even the solubility of a solid is different on different faces of a crystal.

For all these reasons, if a crystal is broken into any irregular shape, it is easy, in spite of this, to identify it as a piece of crystalline matter ; while on the other hand, if a mass of glass is cut into the exact shape of a crystal, it is equally easy to show that it is a fragment of non-crystalline (amorphous) material artificially shaped.

Heat some clear lime-water to boiling in a test-tube. Take the lime-water out of the bottle with a glass tube, and avoid stirring up the material at the bottom. This applies only to lime-water. All ordinary liquids are to be poured carefully from the bottle, not withdrawn by means of tubes. Cork the lime-water bottle at once after using it, since the air spoils it. Does the liquid become

turbid? Does heat affect the solubility of lime in the same way as it does that of potassium chlorate? How do almost all solids behave in this respect?

EXPERIMENT 14.—**Supersaturated solutions.**—Place a few drops of water in a test-tube, half fill the tube with sodium thiosulphate—called "hypo" by the photographer—and heat with a small flame kept in constant motion. When complete solution has occurred, pour the liquid into a clean tube and cork it to exclude dust. Let it cool. Does it behave like potassium chlorate solution on cooling? Throw into the cold liquid a fragment of solid sodium thiosulphate. (?) What is a *supersaturated* solution? Do you regard it as a stable or an unstable state of things? Would *any* crystal answer the purpose, if dropped into the supersaturated solution in this experiment? Since a solution which does not contain any of the undissolved solid may be either unsaturated or supersaturated, what is the only way of being sure that a solution is *saturated?*

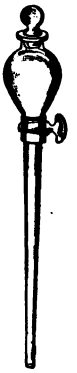

FIG. 14.

EXPERIMENT 15.—**Solutions of liquids in liquids.**—To a test-tube of water add a drop of chloroform. Cover the tube with the thumb and shake it. Are the liquids apparently insoluble in each other? Are they really insoluble in each other?

Half fill a small separatory funnel (Fig. 14) with water. Add a little ether, insert the stopper and shake. *Ether is highly inflammable, and must not be used in the neighborhood of a flame. The ether-bottle must be kept tightly corked.* Add another small quantity of ether and shake again. (?) Now add about half as much ether as there is water present, and shake. Is ether soluble in water? Is there a limit to its solubility? What is the composition of the two layers in the funnel? Give a suitable name to each layer. Allow about 30 c.c. of the lower layer to

run into a 100 c.c. beaker, and get rid of the remainder of the liquid by running it down the sink with an abundance of water. Place the beaker on wire gauze and, having first made sure that no one in the neighborhood is working with ether, heat it gently. Hold a lighted match over the beaker. (?) Did the lower layer contain ether? Is the solubility of ether in water increased or decreased by heating? The solubility of water in ether is increased by heat.

Fill a test-tube one-fourth with alcohol and add water in small quantities, shaking after each addition until the tube is full. What is the result? Is it possible to obtain two layers by mixing alcohol and water? Some other pairs of liquids—for instance, alcohol and ether—behave in the same way.

EXPERIMENT 16.—**Solutions of gases in liquids.**—Heat some faucet water in a beaker on wire gauze not quite to boiling. (?) Heat some ammonia water in a beaker. Hold a burner flame over the beaker. Is there any evidence that a gas escapes? What is the effect of rising temperature on the solubility of a gas?

Fill a 100 c.c. beaker with soda-water from a siphon bottle. What is soda-water? What was the cause of the effervescence when the liquid escaped into the beaker? How does the solubility of a gas vary with the pressure? Lower into the liquid a sealed tube, open end down. Throw into it a fragment of charcoal. Explain the results.

Taste some of the soda-water. Place about 10 c.c. of it in a test-tube and boil it for a time. Cool by running water over the tube and taste again. Is the taste the same? Why? What is the cause of the unpleasant taste of boiled water, and how can it be remedied?

Problems

6. 10.98 grams of a solution of potassium chlorate saturated at 18° was placed in a weighed dish and evaporated to dryness. The residue weighed .7025 gram. How much potassium chlorate was contained in 100 parts of the solution?

7. Taking the figures stated in problem 5, how much potassium chlorate will 100 parts of water at 18° dissolve?

8. Taking the same figures, how much water at 18° is necessary to dissolve 1 gram of potassium chlorate?

9. A solution of common salt saturated at 15° contained 26.39 per cent of salt. How much salt will 100 grams of water dissolve at 15°?

CHAPTER III

PHYSICAL AND CHEMICAL CHANGE—LAW OF THE INDESTRUCTIBILITY OF MATTER

Experiment 17.—**Changes in matter.**—Take a piece of platinum foil in forceps and hold in the Bunsen flame. Let it cool. (?)

Tear up a little paper in very small pieces and bring near it a roll of sulphur. Rub the sulphur briskly on the coat-sleeve and again bring it near the paper. (?)

Crease a long narrow piece of paper in the middle so as to make a trough, which will slip into a test-tube. By means of this introduce a little *mercuric iodide* into a clean, dry test-tube, without getting any of it on the sides. In doing this hold the tube horizontally and slip it over the paper, near the end of which the mercuric iodide is placed. Then upset the trough and deposit the mercuric iodide in the bottom of the tube. Always use this method of introducing powders into tubes when you desire to keep the upper portion of the latter clean.

Lay a cork loosely in the mouth of the tube and weigh

it or balance it with copper filings or iron filings. Before weighing be sure that it is absolutely clean and dry on the outside, and all through the experiment handle the tube with paper—not with the fingers—to avoid soiling it.

Now heat the tube very gently. It should be simply brushed with the flame once in two or three seconds and the powder should be constantly shaken about. The temperature must not rise much above the boiling-point of water. *Overheating will melt the mercuric iodide and spoil the result.*

When the change is complete stop heating, let the tube cool, and weigh it again. Is the weight the same? What does the result show? What is the substance in the tube? Throw out half of it on a piece of paper and rub it with a glass rod. (?) Let the rest of it stand over night and examine it in the morning. (?) What is the natural state of mercuric iodide at ordinary temperatures? At slightly elevated temperatures? In what respects does this change resemble and in what respects differ from the transformation of water into ice? In what respect do all the changes carried out in Experiment 17 resemble each other? Mention some other changes in matter which are like them in this respect.

EXPERIMENT 18.—**Changes in matter. The law of the indestructibility of matter.**—Take a short piece of magnesium ribbon in forceps and hold one end of it in the flame. Receive the product in a dish and examine it. It is called *magnesium oxide.* Why? Compare this change with the heating of platinum.

Place in a small hard glass test-tube enough *mercuric oxide* to fill it to the depth of 1 cm. or more. Introduce the powder by means of a paper trough. Clamp the tube horizontally, placing the clamp near the mouth so that heat can be applied without spoiling the clamp, and heat gently, brushing the tube with a small flame. Hard-glass

tubes must be heated with great caution, for they break very readily. When the mercuric oxide changes color, examine it and let the tube cool. What kind of a change is this? When mercuric acid is cooled to —200° C. it becomes sulphur-yellow, and resumes its original tint on being allowed to warm to room temperature.

Fit the test-tube with a perforated cork bearing a delivery tube, and arrange the apparatus as shown in Fig. 15. Apply heat, at first very gently, gradually raising the

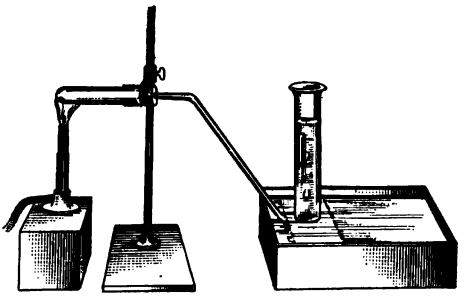

FIG. 15.

temperature to redness. Collect two test-tubes and a small cylinder of the gas given off. Apply the yellow flame of the burner and cover the hot tube with a layer of soot to make it cool gradually and avoid cracking. Disconnect the apparatus. Examine the residue in the tube. What is it? Record the properties of the gas collected. Into a test-tube of it introduce a spark. (?) Slip a glass plate under the cylinder full of the gas and place it upright on the desk. Draw aside the plate for an instant and introduce into the gas a piece of burning candle, held erect on a bent piece of stiff wire. (?) What is the gas? *Keep the flame out of contact with the glass.*

Sketch the apparatus in your note-book and write a full description of your work. What is the historical interest of this experiment? Clean the hard-glass tube when

perfectly cold by washing it out thoroughly with water, then allowing a little strong nitric acid to remain in it for a time and finally washing with water again.

Weigh roughly about 1 gram of potassium iodide and dissolve it in a small test-tube about half full of water. Weigh roughly 0.7 gram of mercuric chloride and dissolve it in about 100 c.c. of water in an Erlenmyer flask. *Mercuric chloride is very corrosive and poisonous, and must not be touched with the fingers. If any of it gets upon the hands it must be removed at once by washing with water.* Carefully place the test-tube in the flask. Remember that chemical glassware is purposely made thin and will not endure violent treatment. Hold the flask upright, see that it is absolutely dry and clean on the outside, and cork it tightly. The completed apparatus is shown in Fig. 16.

FIG. 16.

Now take hold of the flask by means of a folded paper and place it on the trip-scales or a large lecture-table balance. Weigh it accurately, or balance it by metal filings. Record the weight, or leave the filings on the balance; take the flask with the paper holder and tilt it so that the liquids mix. What happens? The red substance is mercuric iodide. This accounts for the mercury and the iodine, but there must be another compound produced. Of what two elements? What becomes of this compound? Devise a method of obtaining it.

Put the flask back on the balance and weigh it again. How do the weights compare? What law have you exemplified? State the law in several different ways. Why is it important? What evidence does astronomy furnish of its truth?

CHAPTER IV

MIXTURE—ELEMENT—COMPOUND—THE LAW OF DEFINITE PROPORTIONS

EXPERIMENT 19.—**Mixtures and compounds.**—Powder finely some roll sulphur (about 3 grams) and mix most of it with about an equal weight of iron filings. Reserve a little. Examine a little of the mixture with a lens. Can you perceive iron and sulphur in it? Place a little of the mixture in a beaker half full of water and stir it. (?) Devise a method of separating the two based upon this behavior.

Place a little iron filings in one dry test-tube, and a little powdered roll sulphur in another, and try the behavior of both with a few drops of *carbon disulphide.* Which dissolves?

Carbon disulphide is inflammable and must not be used in the vicinity of flame. The bottle must be kept corked. The vapor is injurious and should be inhaled as little as possible.

Now place a little of the mixture of iron and sulphur in a dry test-tube, and shake it up with about 5 c.c. of carbon disulphide. Allow to settle, and pour off the clear liquid into a dry dish. What is the substance in the test-tube? What remains in the dish after the carbon disulphide has evaporated?

Place a small portion of the mixture of iron and sulphur on a piece of paper and stir it with a magnet. Explain.

Now transfer the rest of the mixture to a dry test-tube, tap the tube on the table to make a channel along the upper surface, clamp the tube horizontally near the mouth and apply a burner flame to the extreme bottom. It is well to place a tin pan under the tube in case of breakage.

Here, and in all such cases, do not keep the hand directly under the tube while heating. When the reaction begins, remove the flame and observe. Is energy evolved or absorbed when iron and sulphur combine? Let the tube cool, break it, if necessary, and examine the product. It is a chemical compound of iron and sulphur, called *iron sulphide.* Powder it, and examine some of it with a lens. Are iron and sulphur visible in it? Will carbon disulphide extract the sulphur from it?

Explain all the distinctions between a compound and a mixture which are exemplified in this experiment. Notice that all these differences are the result of the fact that the mixture is composed of *two* substances, each having its own properties, while the compound is *one* substance.

However, the greatest distinction between compounds and mixtures does not appear at all in this experiment. What is it?

The exact quantities in which iron and sulphur combine to form iron sulphide is 56 parts of iron to 32 parts of sulphur. Clearly we took more sulphur than was necessary. What became of it? What would have been left in the test-tube if we had taken more iron than was required for the sulphur?

EXPERIMENT 20.—**The law of definite proportions.**—Clean and dry a porcelain crucible carefully and weigh it with the cover. Record the weight. Count the weights at least twice, once while they are upon the pan of the balance and once upon returning them to the box, to be sure there is no mistake. Handle the weights with forceps. They must not be touched with the fingers. Be sure that both pans of the balance are *clean* before weighing.

Introduce into the crucible from 0.3 to 0.5 of a gram of powdered magnesium and weigh again accurately. Be sure that none of the magnesium is upon the balance pans,

since this will make all your work useless. Again count the weights at least twice.

Support the covered crucible on a pipe-stem triangle on a ring of the stand and allow a Bunsen flame to touch the bottom. After ten minutes place the cover so as to allow a very limited access of air. *The operation must be so conducted that nothing escapes from the crucible.* If any white smoke (magnesium oxide) should appear, the crucible must be covered at once. After ten minutes more, lift the cover by the ring with the forceps and shake into the crucible any magnesium oxide which may adhere to it, turn up the flame until it half covers the open crucible, and heat strongly for five minutes, turn down the flame gradually, extinguish it, cover the crucible, let cool five minutes, and weigh. The magnesium oxide in the crucible should be white or grayish-white.

The following will serve as an example of the proper method of record and calculation:

Weight crucible, cover, and magnesium	15.06
" " and cover, empty	14.64
Magnesium taken	.42
Weight crucible, cover, and magnesium oxide	15.34
" " " and magnesium	15.06
Oxygen taken from air	.28

Therefore, 0.42 gram of magnesium combines with 0.28 gram of oxygen. Calculate the oxygen which will combine with 1 gram of magnesium, thus:

$$.42 : .28 = 1 : x.$$
$$x = \tfrac{2}{3} \text{ gram, or } .66.$$

Calculate also the quantity of magnesium which would combine with 1 gram of oxygen, thus:

$$.28 : .42 = 1 : x.$$
$$x = 1.5 \text{ grams.}$$

Calculate in the same way the quantity of oxygen which would combine with 24 grams of magnesium and the quantity of magnesium which would combine with 16 grams of oxygen.

If time permits, repeat the entire experiment. If not, compare your results with those of others who have done the same work. If the experiment is properly performed, different results should be nearly the same. (Why not exactly the same?) If your result varies widely from the figures given above, first look for errors in the calculation, and if none are found, repeat the experiment.

Does the same weight of magnesium always combine with the same weight of oxygen? Is the composition of magnesium oxide always the same? How about other compounds? State the law. What bearing have the results obtained in the electrolysis and synthesis of water upon the law? Explain exactly what is meant by the statement that magnesium and oxygen are elements and magnesium oxide a compound of them. Why are solutions *mixtures* although they are homogeneous and can not be separated by mechanical methods?

CHAPTER V

HYDROGEN

EXPERIMENT 21.—**Reaction of sodium and water.**—*Sodium must not be touched with fingers which are in the least moist either with water or perspiration. Your desk must be dry when working with it and everything with which you touch it must be scrupulously dry. The air acts rapidly upon sodium and it must not be exposed. Take only a small piece from the bottle at a time, and imme-*

diately cork the latter. The liquid over the sodium in the bottle is naphtha or kerosene, and the bottle must not be opened in the vicinity of a flame. No sodium must be put away under the desk or allowed to remain lying about, since it may catch fire. None must be thrown into the waste-jar, since it may ignite the paper or other substances which the jar contains. Return any unused portions to the bottle, or place in a vessel specially provided for that purpose.

Throw a clean piece of sodium, free from crust, half the size of a pea—*no larger*—into a cylinder half full of water. Immediately cover the cylinder with a glass plate or a piece of paper, and *wait until the reaction is over before removing the cover.* The action usually ends with a slight explosion which may endanger the eyes if the cover is removed too soon.

Describe what happens. Throw in another similar piece of sodium. Feel the water of the cylinder between the fingers. (?) Taste a little of it. (?) *Here, and always, immediately reject the liquid tasted and rinse out the mouth with water. Do not taste substances unless directed.* Try the behavior of the liquid with red and blue litmus paper, pieces about 1 cm. square or less. (?) Allow to fall into the liquid a drop of a solution of phenol phthalein. (?) What are these new properties of the liquid due to?

Wrap a clean piece of sodium half the size of a pea in dry tea-lead (lead foil). Punch several holes in the lead with a knife-blade. Invert a test-tube full of water in a tin pan containing water, and quickly slip the lead containing the sodium under it. If necessary use another smaller piece of sodium wrapped in lead to complete the filling of the tube. *Take no more sodium than is directed.* The use of larger quantities is likely to cause explosions which may imperil the sight.

What gas collects in the tube? Does it come from the water or the sodium? Has the gas any color or odor? Is it soluble in water? Does it burn? Quietly or with explosion? To what product?

Record exactly what you observe. For instance, if you find the gas has an odor, record it, and if you think the fact peculiar, inquire of the instructor. *Never alter your observations to correspond to preconceived notions.*

EXPERIMENT 22.—**Hydrogen from zinc and sulphuric acid.**—Place about 20 grams of granulated zinc in a gas-generating bottle. The bottle must be held almost horizontal and the zinc allowed to slide into it, otherwise the shock of the falling zinc will break it. The bottle is provided with a doubly perforated cork carrying a funnel tube and a delivery tube, and the apparatus is arranged for collecting the gas over water (Fig. 10, Part I).

Insert the cork tightly with a twisting motion, and pour in through the funnel tube enough water to cover the zinc thoroughly. The funnel tube must dip into this water. Extinguish any burner flame that may be in the neighborhood, and slowly add strong sulphuric acid through the funnel tube until gas is briskly evolved. Do not add too much acid. The maximum should be about one-fourth as much by volume as there is water present. Allow the gas to escape through the water for three minutes. Why? *Do not attempt at any time in the experiment to light the gas at the exit tube.*

Collect the gas over water in wide-mouthed bottles of about 300 c.c. capacity. Determine its properties. Has it any color or odor? Is it soluble in water? Does the method of collecting it throw any light on this last question? In order to obtain more definite information fill a test-tube half full of hydrogen and mark the level of the water by a strip of gummed label on the tube. Then shake the tube for a time, keeping its mouth under

water. (?) Use this method hereafter in testing the solubility of gases.

Will the gas burn? Try a bottle of it—*not the exit tube of the generator.* Will it support combustion? Try it by holding a bottle-full mouth downward and introducing a lighted candle fastened on a wire. Keep the candle out of contact with the walls of the cylinder so as not to wet the wick. Withdraw the candle slowly. (?) Repeat.

Fill two bottles of the same size with the gas. Support one in an inverted position in a ring of your stand, the mouth not touching the table. Place the other upright. Uncover them at the same moment and allow both to remain uncovered for two minutes. Now thrust a lighted taper or match into each in turn. Draw conclusions.

Fill a small strong bottle over water $\frac{1}{4}$ with air and $\frac{3}{4}$ with hydrogen. Ignite the mixture. Explain the cause of this behavior? Why does not the flame strike back from gas-jets along the mains to the gas-works? Would it be safe to supply cities with a mixture of gas and air by means of pipes?

Procure a strong round-bottomed ginger-ale bottle and a cork which will fit it. Perforate the cork so that an artificial straw like those used at soda-fountains will fit fairly well in the hole. Put in the bottle some granulated zinc and water and add enough sulphuric acid to produce effervescence. Immediately insert the cork bearing the straw and light the end of the straw, holding the bottle upright, and the face *away from over it.* What is the cause of the explosion? Would an explosion have occurred if you had waited several minutes before lighting the straw? Why is it dangerous to light the hydrogen at the exit tube of your generator? *Whenever it is necessary to do this you should wrap the generator, cork, and funnel-tube carefully with a towel, so that no harm can*

result, and wait five minutes before applying the light. This applies not only to hydrogen but to all combustible gases.

Filter the liquid which remains in your gas-generating bottle into a clean porcelain dish. Prepare the filter by

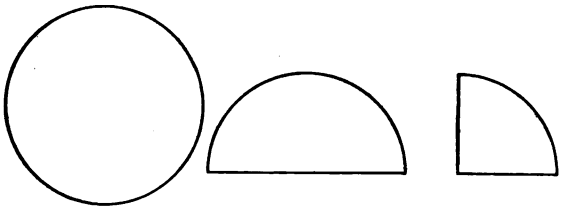

Fig. 17.

folding it first into a semicircle, then into a quadrant (Fig. 17). Thrust it tightly into a funnel and wet it to keep it in place. Support the funnel with its stem against the side of the dish, so that the liquid will slip down the side of the dish and not splash. Pour the liquid into the filter by means of a glass rod to avoid splashing (Fig. 18). Never fill up quite to the edge of the filter, but keep the latter nearly full, as this makes the process more rapid.

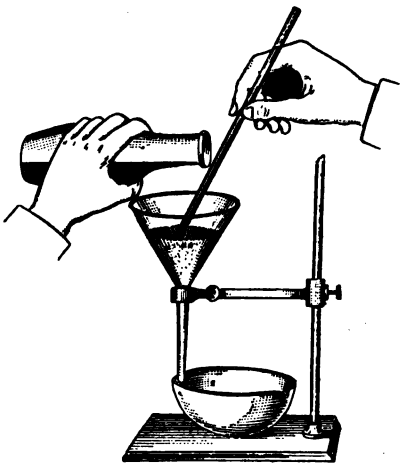

Fig. 18.

When the liquid has run all through wash off the zinc in the bottle, and return it to the stock bottle of granulated zinc. Evaporate the liquid in the dish down to one-fourth its bulk. No wire gauze is needed in

heating the dish. *The flame must not be allowed to touch that portion of the dish which is above the level of the liquid in it, and the dish must be dry on the outside.* The last sentence applies to all heating of liquids in vessels of glass or porcelain. Reduce the size of the flame as the liquid becomes less.

When only one-fourth of the liquid remains, let it cool. If nothing happens on cooling, evaporate carefully one-half of the remaining liquid and allow to cool again. Describe in your notes the crystals which separate. They are called *zinc sulphate,* and they contain zinc, sulphur, and oxygen. Sulphuric acid contains hydrogen, sulphur, and oxygen, so that we may describe the change which takes place in the gas-generating bottle by the statement that the zinc takes the place of the hydrogen.

CHAPTER VI

OXYGEN AND HYDROGEN PEROXIDE

EXPERIMENT 23.—**Preparation and properties of oxygen.**—Place about a gram of potassium chlorate—free from dirt—in a clean, dry test-tube and apply a gentle heat. At first the tube must be simply brushed with the flame. Notice the *decrepitation* and melting of the salt. Apply the spark test. (?) When no more gas escapes examine the residue in the tube. It is potassium *chloride.* Compare its taste with that of potassium chlorate. Heat some of it in a small tube of hard glass, sealed at one end. Does it give off oxygen? Does it melt as readily as potassium chlorate?

In a dry, clean test-tube melt carefully about a gram of potassium chlorate, remove from the flame and show that oxygen is not escaping by the spark. Now throw into

the tube about ½ gram of powdered manganese dioxide and test again with the spark. (?) What is *catalytic action?*

Make a mixture of two parts of potassium chlorate and one part of manganese dioxide, by weight. Both substances should be *coarsely* powdered and *free from dirt.* Fill a wide test-tube to the depth of about 6 cm. with this mixture, make a channel by tapping on the table, and clamp the tube horizontally near the mouth. Provide a well-fitting cork with a delivery tube to collect the gas over water. Brush the tube with the flame. *Rapid heating or the presence of dirt in materials or tube may cause dangerous explosions.* When the gas begins to escape rapidly, remove the flame, and do not begin heating again until the evolution of gas slackens or ceases. Only a very moderate heat is required. A high temperature will melt the tube and spoil the experiment.

Collect the gas in wide-mouthed bottles. Let them remain with mouths under water until wanted, or, if there is not room for this, cover them with wet filter-paper plastered tightly over the mouth. Also collect half a test-tube full of the gas to investigate its solubility in the same way as with hydrogen. If the oxygen supply fails, remove the exhausted material in the test-tube by poking it with the handle of a spoon, put in fresh material and again heat gently. Record the properties of the gas, including its color, odor, and solubility. The cloudiness which it shows at first is due to dust carried over from the test-tube and will disappear on standing. Inhale a little of the gas.

Heat a piece of charcoal in a deflagrating spoon until it glows at one point. Remove it from the flame and plunge it into a bottle of oxygen. (?) Cover the jar with a glass plate and put it aside. Take another piece of charcoal and burn it in a bottle of air in a similar way. It may be necessary to heat the charcoal several times. Cover this bottle also. Examine both bottles. What is

evidently the nature of the product of the burning of charcoal? Into both bottles pour a little lime-water and shake. (?) Cloudiness produced in lime-water is evidence of the presence of *carbon dioxide* (so-called carbonic-acid gas). Does charcoal produce the same substance when it burns in oxygen as when it burns in air? Does it produce the same amount of heat? Why is the burning in oxygen more rapid and brilliant?

Put some sulphur in your spoon and start it burning in the air. Place the burning sulphur in a covered bottle containing air and allow it to burn for a time. Cover the jar and stand it aside. Place the burning sulphur in a bottle of oxygen. (?) Compare the two bottles. What is the nature of the combustion-product of sulphur? Notice *carefully* the odor of the contents of each bottle. The odor is extremely irritating and suffocating. Shake up a little water in each, and add to the water a small piece of red and a small piece of blue litmus paper. (?) Heat the spoon red-hot to remove the sulphur and allow it to cool perfectly before using it again or putting it away.

Straighten a watch-spring by pulling it out into a line and drawing it slowly through the Bunsen flame. Slip aside the cover of a bottle of oxygen an instant and throw in a layer of sand 1 cm. deep. Cut a piece of match-stick 1 cm. long and split it for half its length. Slip this over the end of the watch-spring, set fire to it, and introduce it into a bottle of oxygen. The bottle must be kept covered during the experiment and the spring introduced between bottle and cover. If nothing happens, place another split match-stick on the spring and try again. What is the result? Describe the iron oxide produced.

Put some magnesium powder in a small, clean, cold deflagrating spoon. The powder should project a little above the edge of the spoon. Start its combustion by letting the burner flame play upon it and let it burn in the

air. *Let cool,* and examine the magnesium oxide produced. It is dangerous to remove the substance until perfectly cold. Place some of it on a small piece of red litmus paper and wet it. After a minute remove the magnesium oxide. (?) Now fill the spoon again with magnesium powder, start it burning, and plunge it into a bottle of oxygen. (?) Examine this magnesium oxide and test it with red litmus paper as you did the other. Are the two identical?

Place a clean piece of sodium in your spoon—which must be cold and clean. Heat until the combustion starts and place in a bottle of oxygen, which must be kept covered during the experiment. After the combustion, introduce a little water into the bottle and shake. Test the liquid with red and blue litmus paper. Use very small pieces.

Sulphur is a non-metal. Notice that its oxide combines with water to produce a substance which turns litmus red—a substance which is an acid. Magnesium and sodium are metals. Their oxides combine with water to produce bases—substances which turn litmus blue. This is an important difference between metals and non-metals.

EXPERIMENT 24.—**Action of oxygen upon magnesium and iron.**—Cut a piece of asbestos board about 10 cm. square, and heat it for five minutes, to drive off moisture. Let it cool, place it upon the trip-scales or a lecture-table balance and place upon it a heap of magnesium powder about 2 cm. in diameter. Balance the other pan of the scales with small shot, copper filings or iron filings. Apply a burner flame to the tip of the heap. (?) Let cool and add shot or filings as required to restore equilibrium. Explain exactly what has occurred. Repeat the same experiment, using powdered iron in place of magnesium.

EXPERIMENT 25.—**Quantitative experiment. Decomposition of potassium chlorate. Weight of a liter of oxy-**

gen.—Before the student can make the calculations connected with the following experiment he must understand the methods of correcting the volumes of gases for temperature, pressure, and the presence of water vapor. Unless he has already taken up this subject in his course in physics, he should study the discussion in the Appendix and solve the problems there given. Fit up the apparatus shown in Fig. 19. *T* is a small hard-glass test-tube. *S* a tightly fitting stopper of red rubber, the tube from which just passes through the closely fitting rubber stopper of the acid bottle *A*. This bottle is filled with water which has stood in the laboratory over-night, and is approximately at room temperature. *G* is a liter graduated cylinder. The tube *A A'* is filled with water by placing *A* in a beaker of water and sucking at *A'*. Then the Hoffmann clamp *C* is tightened and the stopper *R* tightly inserted, the tube *T* being removed. This tube is cleaned carefully, brushed with a flame until warm, and accurately weighed. Not more than 2.5 nor less than 1.5 grams of potassium chlorate is introduced into it, and it is brushed again with the flame until the potassium chlorate is just melted. *The heating must be stopped before any oxygen escapes.* Allow the tube to cool, and if any water has appeared in the upper portion, absorb it with a wisp of filter paper. Weigh the tube accurately. Find the quantity of potassium chlorate taken by subtraction.

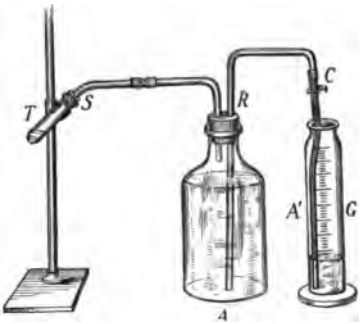

FIG. 19.

Place the cold tube *tightly* on the stopper *S*, loosen the clamp *C* and cautiously heat the potassium chlorate, finally

to faint redness. *The secret of success is gradual heating.* Cool gradually, keeping the clamp C open and the rubber tube A' below the level of the water in G. Disconnect the apparatus. Take the temperature of the water in the bottle A, which is the same as that of the oxygen collected above it. Read the volume of the water in G, which is that of the oxygen (estimate to tenths of the divisions on the cylinder in doing this). Read the barometer in the laboratory, which will give you the atmospheric pressure under which the gas was collected. Repeat all readings to be sure there is no mistake. Record all readings at once in your notebook.

Weigh the tube T accurately and ascertain by subtraction the weight of the oxygen escaped from it. Also find by subtraction the weight of potassium chloride it contains.

Calculation

a. How much oxygen would 1 gram of potassium chlorate yield?

Solved thus:

$$\frac{\text{Weight oxygen escaped}}{\text{Weight potassium chlorate taken}} = x.$$

Has potassium chlorate always the same composition? Does your result agree fairly with those of others who have made the same experiment? Why not exactly?

b. What is the weight of 1 liter of oxygen just as you have collected it—that is, saturated with water and at the temperature and pressure of the air of the laboratory?

Solved thus:

$$\frac{\text{Weight oxygen escaped}}{\text{Volume oxygen in liters}} = x.$$

c. What is the weight of a liter of oxygen under standard conditions—that is, dry at a temperature of 0° and a pressure of 760 mm. of mercury?

Solved in the same way as *c*, but you must first find what the volume of your oxygen would be if dry and at 0° and 760 mm. This you can do, as explained in the Appendix (p. 150), by the formula:

$$V_o = \frac{V \times 273 \times P - w}{(273 + t) \times 760}$$

V_o = standard volume.
t = observed temperature at which your gas was measured.
P = observed pressure.
w = pressure of water vapor at the temperature t (this can be obtained from table in Appendix).

The following will serve as an example of the mode of calculating the results:

Volume of water collected in *G*.....	600 c.c., or .6 liter.
Temperature.....	14° C.
Barometer.....	756 mm.
The weight of the test-tube containing potassium chlorate, dried by being melted, was	12.50 grams.
Weight of empty test-tube.....	10.46
Potassium chlorate taken.....	2.04 grams.
After being heated to redness the tube weighed.....	11.70 grams.
Hence the weight of oxygen which escaped was.....	12.50 grams.
	11.70 grams.
Weight of escaped oxygen =	.80 gram.

a. How much oxygen would 1 gram of potassium chlorate yield? Clearly the answer is equal to

$$\frac{\text{Weight of escaped oxygen}}{\text{Weight of potassium chlorate taken}} = \frac{.80}{2.04} = .39 \text{ gram.}$$

b. What is the weight of 1 liter of oxygen just as collected? The answer is given by

$$\frac{\text{Weight of oxygen}}{\text{Volume of oxygen}} = \frac{.80 \text{ gram}}{.6 \text{ liter}} = 1.33 \text{ grams.}$$

c. What is the weight of 1 liter of oxygen under standard conditions—that is, at 0° C., a pressure of 760 mm., and in perfectly dry condition?

First find the volume of your oxygen at standard conditions. Use the formula

$$V_o = \frac{V \times 273 \times (P - w)}{(273 + t) \times 760}$$

$V = .6$ liter.

$t = 14°$.

$P = 756$ mm.

w (the pressure of water vapor at 14°) we find from table in Appendix = 11.9 mm.

Substituting these values,

$$V_o = \frac{600 \times 273 \times (756 - 11.9)}{287 \times 760} = 559 \text{ c.c., or } .559 \text{ liter.}$$

The weight of 1 liter of dry oxygen at 0° and 760 mm. would be

$$\frac{.80 \text{ gram}}{.559 \text{ liter}} = 1.43 \text{ grams.}$$

Do not expect your results to be identical with those given above. The problem is simply to illustrate the method of calculation, and the numbers were chosen arbitrarily to give results quite close to the truth.

Very careful work has shown that
the weight of a liter of dry oxygen at 0° and 760 mm. is..... 1.429 grams.
And that 1 gram of pure potassium chlorate gives off when heated.. .392 gram of oxygen.

Your results should not vary widely from these figures. If they do the experiment must be repeated; but before repeating it go over the calculation with great care, for the error is more likely to be found there than in the weights or measurements.

EXPERIMENT 26.—**Changes produced in air by respiration.**—Place some clean lime-water in a beaker, blow gently through it by means of a glass tube. (?) Invert three wide-mouthed bottles full of water in your tin pan and collect air in the first bottle from the beginning of an expiration. Withdraw the bottle from the water by means of a glass plate, slip the plate aside an instant and lower into the bottle a lighted candle. Does it burn as long as it would in the same volume of pure air? Why?

In the second bottle collect air from the end of an expiration, using the *last portions* of air from the lungs. Test this with a candle. (?) The result is due partly to the small amount of oxygen, and partly to the large amount of carbon dioxide.

Fill the lungs with air and hold the breath as long as you can without discomfort. Collect the first of the expiration and test it with a candle. Does the result furnish any evidence of the diffusion of carbon dioxide upward, or of oxygen downward, in the lungs?

The body burns up about 220 grams of carbon in twenty-four hours, almost all of which is cast out through the lungs as carbon dioxide. Weigh off roughly this amount of charcoal on the trip-scales in order to get an idea of the quantity. Of course the amount varies greatly in different people and in the same person at different times. The greater the activity and the lower the temperature of the surrounding air, the more active the internal combustion becomes.

EXPERIMENT 27.—**Ozone.**—Set up the apparatus illustrated in Fig. 16, Part I, and perform the experiment

described in the first part of paragraph 54. Make a sketch of your apparatus and write a description of your work. The oxygen required must be obtained from a cylinder of the compressed gas, *not made from potassium chlorate.* If there is no cylinder available, use air instead, but the proof is not conclusive. The induction coil can be excited by two dichromate or three Edison-Lalande cells, or in any other suitable way. Before beginning the experiment read the directions carefully, make a list of the things you need and procure them, so as not to be interrupted at a critical point by the lack of something essential. The solution of starch and potassium iodide must be *very dilute*—not more than 1 gram of starch and .2 gram of potassium iodide in 300 c.c. of water. It should be perfectly clear and limpid.

Procure an ozone tube (Fig. 16, Part I) and prepare ozone as described in paragraph 54, Part I. The apparatus is shown in Fig. 17, Part I. The current of oxygen comes from a cylinder or a gasometer and must be the *gentlest possible.* The slower the current the richer the gas will be in ozone, since it spends a longer time under the influence of the discharge. The preparation of pure ozone by intense cooling is too dangerous and difficult to attempt. Instead of this, note carefully the odor of the issuing gas and hold a piece of starch potassium iodide paper in it.

Place barium dioxide to the depth of about ½ cm. in a test-tube and allow some strong sulphuric acid to trickle down the side of the tube. Test the gas given off with starch potassium iodide paper and notice its odor. Cover the tube with a clean silver coin and stand it in a rack until the end of the laboratory period. If there is no apparent result, do not attribute the fact to an error in your work. The action requires a long time and a gas rich in ozone. Sketch the apparatus in your note-book.

What experimental evidence have you that ozone and oxygen are two forms of the same element? Which is the stable form of oxygen (1) at ordinary temperatures? (2) at slightly elevated temperatures (say 300°)? (3) at very high temperatures?

The existence of the same element in two or more different modifications is called *allotropy.* Do you know of any other elements which exist in *allotropic* modifications? Can you think of any allotropic modifications of carbon? Of tin? When the substance is a *compound* the existence of several modifications is called *isomerism.* What substance have we studied which exists in isomeric modifications?

Rub the head of a match gently with the moistened finger. Notice the odor. (?) Consult the instructor about the preparation of ozone by the action of phosphorus on the oxygen of the air.

EXPERIMENT 28.—**Hydrogen peroxide.**—Examine commercial hydrogen peroxide. Pour a little of it upon some manganese dioxide in a test-tube and test the gas given off with a spark. (?) Take about 25 c.c. of hydrogen peroxide in a beaker and dissolve in it a piece of sodium hydroxide about 1 cm. in length. This is to neutralize the acid which hydrogen peroxide always contains, and which would interfere with the experiments. Add some of this liquid to some water weakly colored with aniline red. (?) Be sure that the water does not contain any solid aniline red at the bottom, for if it does, the solid will continually dissolve and spoil the result.

Put the rest of your hydrogen peroxide into a wide test-tube provided with a cork and a delivery tube leading to a vessel full of water. Boil gently and collect the gas in inverted test-tubes filled with water. Test the gas with the spark. (?) *Only the hydrogen peroxide takes part in the change.* What has happened? What remains in the

test-tube in which the hydrogen peroxide was boiled? Why is the cork frequently driven out of the stock-bottle of hydrogen peroxide?

State *precisely* the relation in composition between hydrogen peroxide and water. State *precisely* what is meant by catalytic action, using hydrogen peroxide as an illustration.

Problems

10. If 75 c.c. of oxygen could be transformed completely into ozone, what volume of ozone would be obtained?

11. 115 c.c. of oxygen were partly converted into ozone. The volume contracted to 110 c.c., but when the gas was gently heated by a burner flame the original volume was exactly restored. Calculate (*a*) the volume of ozone which had been produced and (*b*) its percentage by volume in the gas.

12. 160 c.c. of oxygen containing ozone were heated. The volume became 170 c.c. (*a*) How many c.c. of ozone and (*b*) what percentage of it by volume were present?

CHAPTER VII

COMBUSTION

Experiment 29.—**Combustion.**—Fit up a lamp chimney with a perforated rubber cork and tube for supplying gas (Fig. 19, p. 52, Part I). Cover the chimney with a piece of asbestos having a hole in it large enough to admit your spoon. Clamp the chimney at its upper portion and see that the cork is tightly placed in the bottom. Use illuminating gas—not hydrogen—and make the experiments on combustion described in Chapter VII, paragraphs 62 and 63, Part I.

In the experiment with nitric acid use *fuming nitric acid.* The ordinary acid is not strong enough for the pur-

pose. Keep the burning substances out of contact with the glass to avoid breakage. Always hold your spoon by the upper end, and keep the hand some distance above the chimney, to avoid burning the fingers.

In getting the oxygen flame (Fig. 20, Part I) use a mouth blowpipe the tip of which has been carefully bent until it is parallel with the longer portion. Connect this with an oxygen cylinder and obtain the gentlest current of oxygen possible. Test the strength of the current by letting it blow on the moistened finger. Insert the blowpipe through the flame into the chimney. *Keep the oxygen flame away from the glass.*

The same chimney can be used for the air-flame; but you will need a doubly perforated cork, one end of which carries a wide, straight tube open at both ends (Fig. 21, p. 55, Part I).

Hold a small crystal of each of the four chlorates mentioned in paragraph 63, Part I, in the lower part of the Bunsen flame with forceps. Record the flame colors of the corresponding metals.

What is a flame? Why do some substances burn with flame and others not? Explain what is meant by the statement that the combustion of one gas in another is reversible. Explain exactly what is meant by the statement that the air flame is the gas flame *turned inside out.*

EXPERIMENT 30.—**The slow combustion of iron.**—Stuff a wad of steel wool about half-way up a tube closed at one end and graduated. Wet the steel and clamp the tube, open end down, in a vessel of water. The level of the water inside the tube must be within the graduations. If it is not, remove the tube from the vessel, pour some more water in the tube, cover the end with the thumb, put it again in the vessel of water and clamp it. Read the level of the water inside the tube, allow to stand for half an hour or as long a time as possible, and read the level

again. Examine the steel carefully. What does the result show? Predict the composition of rust. Was heat evolved or absorbed in this experiment? Steel is chiefly composed of iron, and it was the iron alone which was affected in the experiment. Dislodge the steel wool with the handle of a spoon, or with a glass rod, and clean the tube with a small quantity of strong hydrochloric acid.

CHAPTER VIII

No experiments.

Problems

Calculate the percentage composition of the substances whose formulas are given below. The amount of each constituent must be obtained by an independent calculation, never by subtracting the sum of the others from 100. State the results to two decimal places. If the third decimal place is greater than 5 add 1 to the second; if less than 5 discard it.

13. Mercuric oxide..... HgO
14. Water..... H_2O
15. Mercuric chloride..... $HgCl_2$
16. Mercurous chloride..... $HgCl$
17. Potassium chlorate..... $KClO_3$
18. Table salt (sodium chloride)..... $NaCl$
19. Manganese dioxide..... MnO_2

Calculate the simplest formulas of the following substances from their percentage compositions.

20.	Hydrogen..	2.74 per cent.	23.	Mercury....	44.07 per cent.
	Chlorine ...	97.26 "		Iodine....	55.93 "
21.	Nitrogen...	30.43 "	24.	Calcium....	40 "
	Oxygen.....	69.57 "		Carbon.....	12 "
22.	Carbon.....	40.00 "		Oxygen	48 "
	Hydrogen..	6.67 "	25.	Potassium..	52.35 "
	Oxygen....	53.33 "		Chlorine....	47.65 "

CHAPTER IX

SALT AND SODIUM

Remember the precautions necessary in using sodium (p. 28). *Carry out this experiment under the hood. Avoid inhaling chlorine.*

EXPERIMENT 31.—**Salt and sodium.**—Fit up the apparatus shown in Fig. 20, and carry out the *synthesis of salt.* The flask should be small. It is one-fourth filled with manganese dioxide in small lumps—not powder. This must be allowed to slide slowly into the flask, the latter being in an inclined position. *The perforated cork and exit tube must fit tightly;* otherwise chlorine will escape and cause great distress and possibly injury. The bulb must be of hard glass. It contains a fragment of sodium the size of a pea, which may be elongated by rolling, to get it into the tube.

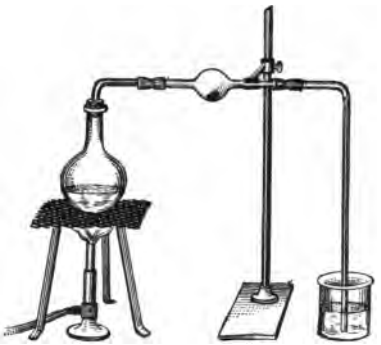

FIG. 20.

Cover the manganese dioxide with strong hydrochloric acid, place the flask on wire gauze, connect the apparatus, and apply a gentle heat. The tube which is bent downward dips into about 20 c.c. of alcohol or sodium hydroxide solution, which will absorb the excess of chlorine.

When the gas in the bulb appears green (?) heat the sodium very gently with a small flame. When the action begins, stop heating the sodium. When the action is over, extinguish the burner, take out the cork of the chlorine generator and fill the latter with water to stop the evolu-

tion of gas. It is well to do this quickly, to hold the breath during the operation, and to retire ten feet or more for a short time until the chlorine which the water forces out of the flask is dissipated.

Break open the bulb. The black product is *silicon* due to the unavoidable action of the sodium upon the glass. Taste the white product. Is it soluble in water? Is it possible for an insoluble substance to have any taste? Why? What information does this experiment give regarding the composition of salt? Is the information qualitative or quantitative? Throw the fragments of the bulb-tube into a vessel of water to get rid of any sodium which may have escaped the action of the chlorine.

Does salt contain more or less energy than its constituents taken separately? How do you know? If you had to decompose salt into sodium and chlorine, would it be necessary to supply energy or would the decomposition furnish energy which could be applied to other purposes? Recall all the cases in your past work in which you have brought about or witnessed the *combination* of two elements. Was energy absorbed or given out in these experiments? Recall the cases in which a compound has been decomposed. Was energy given out or absorbed? Irritation from chlorine can be relieved by at once inhaling the vapor of alcohol.

EXPERIMENT 32.—**Preparation of pure salt.**—Prepare a saturated solution of salt, filter the liquid, and pass hydrochloric acid-gas into it through an inverted funnel, whose rim just dips into the liquid. The gas is obtained as described on p. 56, Experiment 41, the sulphuric acid being allowed to drop *very slowly* into the hydrochloric. Salt, being insoluble in hydrochloric acid, is precipitated; the other substances present remain dissolved. When enough salt has collected, pour away the liquid and wash the salt three times with very small quantities of dis-

tilled water. (Why is this necessary?) Place the salt in a clean dish and apply heat with a small flame which does not touch the dish, stirring the salt constantly with a glass rod. Preserve the pure salt in a dry, clean test-tube, tightly corked and labelled.

EXPERIMENT 33.—**Sodium.**—Examine sodium, and write a description of it. Cut it. Is it a metal? Why does the luster fade? How could the fading be prevented? Could sodium be used in the same way as iron or copper? Why? How does its density compare with that of familiar metals?

Hold a small fragment of sodium in the flame in forceps. (?) Explain, from previous work, the behavior of sodium with oxygen, with chlorine, and with water. Explain *precisely* what is meant, both qualitatively and quantitatively by the equation—

$$Na + H_2O = NaOH + H$$

EXPERIMENT 34.—**Sodium compounds.**—*Sodium hydroxide must not be touched with the fingers. The bottle containing it must be kept tightly corked.* Examine sodium hydroxide, sodium nitrate, sodium carbonate (anhydrous), sodium carbonate (crystallized), and sodium sulphate. Are they soluble in water? What color does each communicate to the flame? Use a clean iron wire hot enough to make a fragment of the sodium compound adhere to it. After the test, clean the wire by dipping it into a little hydrochloric acid in a test-tube (*not in the bottle*) and holding it in the flame until the latter shows no color, and use it to test the next substance.

Prepare a solution of sodium hydroxide by dissolving a piece about 1 cm. long in 50 c.c. of water in a beaker. Test small separate portions of it with blue litmus paper, with red litmus paper (use small pieces), with a drop of phenol phthalein solution, with a drop of cochineal solu-

tion. What are the results? Liquids which affect these coloring matters in this way are called *alkaline.*

Place the rest of the sodium hydroxide solution in a clean dish and add hydrochloric acid to it one drop at a time. After each drop stir with a glass rod and test the liquid with a fragment of blue litmus paper. When the paper is just turned red, remove it and evaporate the liquid carefully to dryness. Turn down the flame very much toward the end and keep it moving. Taste the residue when it is perfectly dry. What is it? What other product must have been formed? Write the equation.

Expose a clean crystal of sodium carbonate and a piece of sodium hydroxide on two separate watch glasses to the air for an hour or overnight. What happens? What is meant by efflorescence? by deliquescence? What has the state of the air to do with both? How would any soluble salt behave if kept in an atmosphere saturated with water-vapor?

EXPERIMENT 35.—**Sodium carbonate. Sodium acid carbonate.**—Stir up some ammonia water with salt in a beaker until the solution is saturated. Let the undissolved salt settle and pour off the solution into another beaker. Pass carbon dioxide into it through a funnel the rim of which dips into the liquid. The carbon dioxide can be obtained from a Kipp generator or made in a gas-generating bottle from marble and dilute hydrochloric acid as directed on p. 133, Experiment 127. The precipitate is sodium acid carbonate (baking soda), $NaHCO_3$. Filter and dry the product by pressing it between layers of filter paper, which are renewed as they become wet.

Arrange a test-tube with a cork and a delivery-tube dipping into a little lime-water. Place some dry sodium acid carbonate in it, support the tube horizontally, and apply a very gentle heat. What gas is given off? What condenses in the tube? Compare the residue in the tube

with anhydrous sodium carbonate, Na_2CO_3. Write the equation.

These two reactions illustrate the *ammonia soda process,* the most important method of manufacturing sodium carbonate.

Make a mixture of sodium acid carbonate with about twice its weight of cream of tartar. Grind the two together in a dry mortar. What familiar substance have you made? Is there any action? As a rule, chemical action between solids at ordinary temperature and pressure is slow, so slow that it is almost imperceptible. Place some of the mixture in a test-tube and add water. (?) Fit a cork bearing a delivery tube to the test-tube and pass some of the gas into lime-water. (?) What is the function of this gas in the baking process? The same gas is produced when sodium acid carbonate is treated with an acid—e. g., hydrochloric acid. (Try it.)

EXPERIMENT 36.—**Water of crystallization.**—Examine some crystallized copper sulphate. Place about 2 grams of it in a dry, clean dish, cover with a dry, clean beaker, and heat by a small flame, kept in motion. (?) After about five minutes, remove the beaker and heat the substance a little more intensely, stirring it and pressing the lumps gently with a glass rod. Let the product cool and compare it with the original substance. What is it? Sprinkle some water upon it. (?) It becomes hot. What is the source of the heat? What is water of crystallization and why is it so called? From the facts furnished by this experiment, discuss the question whether water of crystallization is chemically combined or merely mixed.

PROBLEMS

26. Calculate the percentage composition of

a. **Sodium carbonate..... Na_2CO_3,**

b. **Cystallized sodium carbonate..... Na_2CO_3, $10H_2O$.**

In *b* calculate *water*, not hydrogen.

27. How much sodium carbonate can be made from 10 kilos of salt?

28. How much sodium is necessary to decompose 36 grams of water, and how much hydrogen is liberated?

29. How much sodium can be obtained by the electrolysis of 20 grams of sodium hydroxide?

30. A piece of sodium was placed in water. 500 c.c. of hydrogen under standard conditions escaped. What was the weight of the sodium?

31. How much sodium carbonate and slaked lime are needed to make 2 kilos of sodium hydroxide?

CHAPTER X

CHLORINE

EXPERIMENT 37.—**Chlorine.**—*Carry out this experiment under the hood. Avoid inhaling the gas. Be sure that the cork of the apparatus fits tightly. If the throat becomes irritated, place some alcohol in a beaker or on a handkerchief and inhale the vapor.*

Fit up the apparatus shown in Fig. 25, Part I, and prepare chlorine. Use a small flask ($\frac{1}{2}$ liter capacity or less), 25 or 30 grams of manganese dioxide (coarse powder or lumps—fine powder will not answer), and enough commercial hydrochloric acid to cover it. Proceed exactly as directed on p. 75, Part I, and in Experiment 31, Part II. Collect by downward displacement four wide test-tubes and four bottles full of the gas. Judge when the vessel is full by the color, and *immediately* cover it tightly with a glass plate and substitute another; otherwise the excess of chlorine will be forced out at the top and make the air unfit to breathe. Test-tubes full of chlorine can be corked. The exit tube of the chlorine apparatus must reach to the bottom of the vessel. This tube should be cut about a foot

from the desk and united again by a short piece of rubber tube, that it may be moved without disturbing the apparatus. The vessel in which you are collecting chlorine should be kept covered, and the exit tube slipped between the cover and the side of the bottle. The bottles of chlorine must be kept covered during all experiments. When you have finished and desire to get rid of the chlorine, place the bottles under the hood, remove the covers without breathing, and at once retire to a distance. After the chlorine has escaped—which will require ten minutes—the bottles can be cleaned.

Record the physical properties of the element. Invert a test-tube full in water and shake gently. Is it soluble? Into a test-tube filled with the gas throw a pinch of powdered arsenic. (?) Into another test-tube a little powdered antimony. (?) Fill a *small,* clean test-tube with hydrogen from the Kipp apparatus by upward displacement. Without losing any of either gas, bring this tube mouth to mouth with a test-tube filled with chlorine. The two tubes should be of the same size, and the tube containing the hydrogen uppermost. Keeping the mouths together, invert several times to mix the gases, then hold the mouth of each tube to a flame. This experiment must not be made in direct sunlight. Why not? What is the product?

Lower a hydrogen flame into a bottle of chlorine. If you use a gas-generating bottle in this experiment, remember that it must be allowed to run five minutes and that the generator and stopper must be wrapped in a towel before lighting the hydrogen. Describe the appearance of the flame of hydrogen burning in chlorine. Hold a glass rod bearing a drop of ammonia near an open bottle of hydrochloric acid. (?) This is a *test* for hydrochloric acid. Now hold a drop of ammonia in the gas left in the jar in which the hydrogen was burned. Is it hydrochloric

acid? Has it the color of chlorine? Heat a piece of charcoal in a spoon until it begins to burn, and place it in chlorine. (?)

Take out the charcoal and lower a lighted candle supported on a wire into the same bottle of chlorine. The candle is composed chiefly of compounds of hydrogen and carbon. The black substance thrown off from the flame is carbon (soot). Use the two preceding results to explain this one. What else must be produced? Let the covered jar stand till the soot settles and try the test with the glass rod bearing a drop of ammonia.

In the third bottle place a bit of red litmus paper, a piece of blue litmus paper, a strip of colored calico, and a fragment of printed matter smeared over with writing ink until it is illegible. Describe and explain the results. What is the most important use of chlorine?

In the fourth bottle place some fresh slaked lime (milk of lime answers well) and immediately cover the bottle tightly with the palm of the hand and shake it. (?) Does the chlorine disappear?

Carefully add nitric acid, a few drops at a time, to the contents of the bottle. What happens? What very important technical process does this illustrate?

EXPERIMENT 38.—**Chlorine water. Bleaching.**—Prepare some chlorine water by passing chlorine made as in Experiment 37 into cold water under the hood for twenty minutes. You will need about 500 c.c. Examine it. Has it any of the properties of the gas? Does it bleach? Try small portions of it with red and blue litmus paper. With a drop of ink.

Select a tube of soft glass 1 cm. or more wide and about 1 meter long. Seal it near one end by drawing it out in the flame of the blast lamp. Cover it with soot with the yellow flame before cooling, to avoid cracking. When cool wash off the soot, fill the tube with chlorine

water and invert it in the same liquid in a glass or porcelain dish. Clamp vertically and allow to stand in direct sunlight as long as possible.

When a sufficient quantity of gas has collected, cover the end of the tube with the thumb, remove it from the dish and test the gas with the spark. Explain what has happened.

Prepare a thin paste of bleaching powder and water in a small beaker or dish. Add a few drops of sulphuric acid (why?) and soak in the liquid a small piece of some colored cotton fabric (calico) or a piece of litmus paper.

EXPERIMENT 39.—**Combustion of chlorine in hydrogen.** (*Use the hood in this experiment.*)—Generate chlorine from coarsely powdered manganese dioxide and strong hydrochloric acid in a wide test-tube, which must not be more than one-third filled with the mixture. The tube is clamped upright and is closed by a perforated cork bearing a straight narrow tube about 20 cm. long ending in a jet. Have ready a bottle of hydrogen, standing in water. Apply a gentle heat to the chlorine generator, and when the gas over the liquid is green and chlorine is escaping freely, light the hydrogen, holding the bottle inverted, and instantly slip the bottle over the jet of chlorine. Taken in connection with the burning of hydrogen in chlorine, what does the result show? Test the product in the bottle with a drop of ammonia on the end of a glass rod.

Fill the test-tube with cold water to stop the production of chlorine.

Sum up all the evidence you have obtained thus far bearing upon the composition of hydrochloric acid.

EXPERIMENT 40.—**Electrolysis of hydrochloric acid.**—Place about 200 c.c. of strong hydrochloric acid in a beaker and dilute it with about $\frac{1}{4}$ of its volume of water. Add a handful of salt and stir for five minutes to saturate the liquid. Pour off the clear liquid and introduce it into

the apparatus shown in Fig. 2, Part I, which you have already used for the electrolysis of water. The salt takes no part in the change. Its object is to reduce the solubility of the chlorine in the water present.

Allow the current to pass. Allow some of the gases to collect. Notice the color, and, cautiously, the odor of both. Remember that a colored gas may appear colorless in a thin layer. Look *along* the tube as well as through it. Try to burn each gas, using a very short rubber tube with a short jet, as in the electrolysis of water. Try the action of both gases upon paper wet with a solution containing potassium iodide and starch. After the apparatus has been running about half an hour, let the gases collect for a time and measure the quantity of each. Is the result exactly what you expect? If not, explain.

EXPERIMENT 41.—**Preparation and properties of hydrochloric acid.**—Prepare hydrochloric acid from commercial hydrochloric acid solution and strong sulphuric acid as directed on p. 81, Part I. A flask closed by a doubly perforated cork carrying a separating funnel and a delivery tube is sufficient. The sulphuric acid must be allowed to drop slowly into the hydrochloric, and the delivery tube must be *dry*. Collect by downward dry displacement in bottles, which must be perfectly dry. (Why?) Keep the bottles covered during the collecting and afterward.

What are the color and odor of the gas? Is it soluble in water? Bring the mouth of a bottle of the gas near the surface of water, remove the glass plate and then plunge the bottle into the water. Do not remove the plate under the water. The sudden inrush of water is likely to break the plate and cut the hand. Wrap the bulb of a thermometer in damp filter paper and introduce it into a bottle of hydrochloric acid. Is there any alteration in temperature? Explain. Does the gas burn?

Does it support combustion? Lower a lighted candle on a wire into a bottle of it. In another bottle place a strip of red and a strip of blue litmus paper. (?) Into another bottle throw a wad of filter paper wet with ammonia. (?)

EXPERIMENT 42.—**Preparation of sodium amalgam.**—Weigh out roughly on the trip-scales in a dry beaker about 200 grams of mercury. Weigh also 1.5–2 grams of sodium and cut it into pieces about half the size of a pea. Place the mercury in a clean dry mortar, add one piece of sodium and press it under the mercury with the pestle until there is a flash and a little smoke, showing that the reaction has occurred. Mercury vapor is poisonous, and it is well therefore to do this under the hood. Continue in this way until all the sodium is introduced. The amalgam should still be liquid. If it is solid, too much sodium has been used and it is necessary to add more mercury and mix thoroughly with the pestle. Sodium amalgam, like sodium, is acted upon by moist air and must be preserved in a tightly closed bottle. Its chemical action is simply that of the sodium it contains, but the presence of the mercury makes the reactions less violent.

EXPERIMENT 43.—**Analysis of hydrochloric acid.**—Allow 10 c.c. of water from a pipette or burette to run into a small clean test-tube and mark the level of the water by a narrow strip of gummed label. Dry the tube thoroughly with paper or cloth, not with the flame. Select a straight graduated tube of about 50 c.c. capacity sealed at one end. The portion nearest the open end will not be graduated. Ascertain its capacity by taking the length from the end of the graduation to the end of the tube in compasses or with a strip of paper and transferring it to the scale on the tube. Add this to the graduated portion to ascertain the total capacity of the tube.

The tube must be perfectly dry. Fill it by downward displacement with hydrochloric-acid gas, generated as in

Experiment 41, and dried by passing through a U-shaped tube containing bits of broken glass tubing wet with strong sulphuric acid. Since the presence of air will spoil the result, the generator must be run until all air is expelled from it and the graduated tube—five minutes or more. The use of the hood will save discomfort. The exit tube must run to the bottom of the graduated tube and must be withdrawn gradually while the generator is still running, otherwise it would leave a vacancy which would be filled with air. Have at hand 10 c.c. of sodium amalgam in the measuring test-tube and *at once* introduce it and cover the tube tightly with the thumb. Let the sodium amalgam fall the length of the tube a dozen times or more to make the action complete, bring the open end of the tube under water in a cylinder, remove the thumb, lower the tube until the level of the water is the same inside and out, and clamp it. Allow to stand five minutes and read the volume of the gas.

Cover the tube with the thumb, remove it from the water, invert it and apply a flame to the residual gas. (?) Calculate thus:

Volume of tube to end of graduated portion	50 c.c.
Capacity of ungraduated portion.....	12 "
Total capacity.....	62 "
Since the 10 c.c. of sodium amalgam introduced expel 10 c.c. of gas, the volume of hydrochloric acid taken was	52 c.c.
The tube at the end of the experiment contained	26.4 "

Hence the proportion of hydrogen by volume contained in hydrochloric acid is $\frac{26.4}{52}$, or about $\frac{1}{2}$.

EXPERIMENT 44.—**Action of sulphuric acid on salt.**—Place a handful of salt in a wide-mouth bottle and allow a little strong sulphuric acid to drop upon it. Cover the

bottle. Is heat produced? Sink a lighted candle in the bottle. (?) Blow across the mouth of it. Hydrochloric acid causes fumes in moist air, because it causes the water-vapor to condense, forming a solution of hydrochloric acid. Hold a drop of ammonia in the bottle. (?)

What substance would remain in the bottle if the action of the sulphuric acid was complete? Write the equation. Place about 25 c.c. of water in a beaker. Stand the beaker in a pan of cold water and slowly pour in 50 c.c. strong sulphuric acid stirring constantly. Let the liquid cool *completely* and pour it slowly upon about 30 grams of salt in a small flask provided with a well-fitting cork and delivery tube. This tube must be perfectly *dry,* for any water will absorb the gas. Apply a gentle heat. Be careful not to let the liquid boil over. Collect the hydrochloric-acid gas in dry test-tubes and bottles. The delivery tube must reach to the bottom of the bottle. Is it soluble in water? Try it by placing a test-tube filled with it mouth downward in water. How does it affect red and blue litmus paper? Does it burn? Does it support the combustion of a candle? Pour some ammonia-water on a wad of filter paper and throw it into a bottle of the gas. (?) What is left in the generating flask?

Pour the material in the flask down the sink, running abundant water at the same time, and again put in salt and dilute sulphuric acid in exactly the same way. Pass the gas into a small inverted funnel whose rim dips into water in a dish and whose stem is connected by a rubber tube with the delivery tube. Continue until the materials are exhausted. Use the hydrochloric acid prepared in this way for Experiment 45.

EXPERIMENT 45.—**Action of hydrochloric acid upon metals.**—Try the action of hydrochloric acid upon small quantities of zinc, magnesium, aluminium, and copper. Place the metal in a test-tube, cover it with water, and

gradually add hydrochloric acid solution. If gas is produced test it with a flame. (?) What is the other product of each case? Try the action of hydrochloric acid upon a fragment of marble.

Questions

Answer these questions as briefly as possible. If you know the equation for the change you desire to explain that alone will suffice, but *never invent equations or formulas.* Remember that a formula is a faithful description of the composition of a real compound. Mere arbitrary collections of symbols are worse than meaningless—they are absurd. Never write an equation in your notes unless you are sure, first, that every formula it contains is that of a real substance ; second, that the number of symbols of each kind is the same on both sides of the sign of equality ; third, that the equation describes a process which you know from laboratory or lecture-table experiments to be a real chemical change. Unless the equation satisfies these requirements, discard it, and describe the change briefly in words. Some of the questions can be answered in a number of different ways, all of which are correct.

1. How could you convert the hydrogen of hydrochloric acid into water ?

Answer

$$Zn + 2HCl = ZnCl_2 + H_2.$$
$$H_2 + O = H_2O.$$

In words, the answer would be : Zinc will liberate the hydrogen, forming zinc chloride. The hydrogen can then be burned in the air to water.

2. How would you make Na_2SO_4 from NaCl ?

3. How could you prepare water, taking the hydrogen from sulphuric acid and the oxygen from copper oxide ?

4. Prepare water, taking the hydrogen from sulphuric acid and the oxygen from potassium chlorate.

5. Prepare arsenic chloride, taking the arsenic in the free state and the chlorine from sodium chloride.

6. How could you convert the oxygen of potassium chlorate into ozone ?

7. Obtain hydrochloric acid, taking the hydrogen from water.

8. Convert the oxygen of water into copper oxide.

Problems

32. What weight of chlorine can be obtained by heating 12.5 grams of manganese dioxide with an excess of hydrochloric acid?

33. How much manganese dioxide is needed to make 25 grams of chlorine from hydrochloric acid?

34. How much aqueous hydrochloric acid containing 20 per cent HCl is required to liberate 100 grams of chlorine with manganese dioxide?

35. 20 c.c. of chlorine were mixed with 16 c.c. of hydrogen and the mixture exploded. What volumes of what gases remained in the vessel?

36. Hydrogen was burned in chlorine and the hydrochloric acid produced collected. It weighed 146 grams. What weights of both gases had been consumed?

37. 50 grams of sodium hydroxide are dissolved in water and the solution mixed with a solution containing 50 grams of pure hydrochloric acid. What substances, and how much, will be contained in the resulting liquid?

38. 32.75 grams of zinc are dissolved in hydrochloric acid. What weights of zinc chloride and hydrogen are produced?

39. 40 grams of magnesium are allowed to dissolve in hydrochloric acid. Calculate the weights of hydrogen and magnesium chloride produced.

CHAPTER XI

CHLORIDES—COMPOUNDS OF CHLORINE CONTAINING OXYGEN

Experiment 46.—**Preparation of chlorides.**—Place about 1 gram of zinc oxide in a test-tube, cover with water, and add hydrochloric acid. If necessary, heat gently, but not to boiling. What is produced? Is gas evolved? What becomes of the hydrogen of the acid? Repeat, using magnesium oxide. Write the equations.

Make a strong solution of sodium hydroxide and slow-

ly add an excess of hydrochloric acid, stirring constantly. Write the equation.

Examine a fragment of marble. Coarsely powder a little. Is it soluble in water? Add some hydrochloric acid to the water. (?) When the action is over, evaporate the liquid to dryness in a dish. Is the residue soluble in water? Is it marble? Write the equation.

State four methods by which the salt of a given metal and a given acid can be prepared, and give an example of each.

EXPERIMENT 47.—**Insoluble and slightly soluble chlorides.**—To a dilute solution of silver nitrate, $AgNO_3$, add a little hydrochloric acid. The precipitate is silver chloride, AgCl. Write the equation. Divide the liquid containing the precipitate into three parts. Expose one portion to sunlight, or the brightest light attainable. With the second ascertain whether silver chloride is soluble in nitric acid. With the third investigate its solubility in ammonia.

Mix a few drops of silver nitrate solution with some solution of sodium chloride. What is the precipitate? Prove it. Any soluble chloride would produce the same effect upon silver nitrate.

Precipitate a little solution of *mercurous nitrate,* $Hg_2(NO_3)_2$, with hydrochloric acid. The precipitate is mercurous chloride, Hg_2Cl_2. Divide the liquid containing the precipitate into two parts. With one ascertain the action of strong light. Expose for some time. Treat the other with ammonia. What are the results? How can mercurous chloride be distinguished from silver chloride?

Make a solution of lead nitrate, $Pb(NO_3)_2$, and add hydrochloric acid to it. The precipitate is lead chloride, $PbCl_2$. Write the equation. Heat the liquid containing the lead chloride to boiling and stand it aside to cool. Is lead chloride soluble in cold water? In hot water? How

could lead chloride be separated from silver chloride? From mercurous chloride? Devise a method of separating a mixture of the three chlorides.

EXPERIMENT 48.—**Chlorine peroxide.**—Make the experiments described in paragraph 112, Part I. Use a cylinder of not more than 200 c.c. capacity and *not more* than 0.5 gram of finely powdered potassium chlorate. Cover the bottom of the cylinder with sulphuric acid and introduce the chlorate gradually.

EXPERIMENT 49.—**Potassium chlorate.** *Potassium hydroxide must not be touched with the hands. Use paper or forceps in handling it. Make this experiment under the hood.*—Dissolve about 40 grams of potassium hydroxide in about 100 c.c. of water in a dish and pass chlorine into the liquid through an inverted funnel whose rim dips under the liquid. From time to time test the liquid with a strip of red litmus paper, and when the latter is no longer turned blue (?) stop the chlorine. Pour off the liquid from the white crystals in the dish and dissolve the latter in the smallest possible quantity of hot water. If anything remains undissolved, remove it by filtering the hot liquid, then let the liquid cool. Collect the product which separates on a filter, and dry it between layers of filter paper. Examine it and compare its appearance with that of potassium chlorate. What is the effect of heat upon it? Use a tube sealed at one end and test the gas given off by the spark. Unless your substance was perfectly dry, water will appear in the sealed tube, but it is not a product of the chemical change and can be disregarded.

What is the action of sulphuric acid on the substance you have made? *Use very small quantities.*

How is potassium chlorate made at present, and what are its chief uses?

Questions

1. Why is no gas given off when an oxide or a hydroxide is treated with hydrochloric acid?

2. What is the real meaning of the expression "insoluble"?

3. How could sodium chloride be converted into sodium nitrate?

4. Does chlorine peroxide, ClO_2, contain more or less energy than chlorine and oxygen separately? How do you know?

5. Construct, in the form of a table, a brief comparison of the four elements already studied.

Problems

40. Calculate the percentage composition of bleaching powder, $CaOCl_2$.

41. Determine the name and formula of a compound having the following composition:

Sodium	21.60 per cent.
Chlorine	33.33 "
Oxygen	45.07 "

42. 25 grams of pure marble are dissolved in hydrochloric acid. What are the products and how much of each is produced. How much HCl is consumed?

CHAPTER XII

No experiments.

Questions

1. What are the two possibilities regarding the nature of matter?

2. What is a molecule? Is a molecule a thing or an idea? Why?

3. What is the evidence which causes us to consider the molecule to be composed of smaller particles?

4. Explain precisely what is meant by the term *atomic weight.*

5. State the chemical laws in the language of the atomic theory.

6. Suppose that some one should succeed in transforming lead

into gold – a highly improbable supposition : what changes in the atomic theory would have to be made to explain the discovery ?

CHAPTER XIII

THE ATMOSPHERE—NITROGEN

EXPERIMENT 49*a*.—**Analysis of the air by means of phosphorus.**—*Phosphorus catches fire spontaneously, and to leave it lying around or put away any of it under the desk would be likely to cause the destruction of the building. Burns made with it are poisoned wounds which heal with great difficulty. It must be kept, cut, and handled under water, and the hands must be freed from it before they are brought into the air. Be careful not to get it behind the nails. If it should catch fire while you are working with it pour water upon it.*

The method is to absorb the nitrogen from a measured volume of air, confined over water, by means of phosphorus. The residual gas, consisting chiefly of nitrogen, is again measured. The experiment is shown in Fig. 28, Part I. The water used must have stood in the laboratory a day or two so as to acquire the temperature of the room. The jar should be wide, so as to allow freedom to the hand.

Fill the graduated tube with water, close the end with the thumb and put it, open end down, into the jar of water. Catch hold of the upper part of the tube with a paper holder—*not with the hand*—and allow 20 c.c. to 30 c.c. of air to enter the tube. Hold the tube so that the level of the water inside and outside is the same, and measure the volume of air. Repeat the measurement to be sure there is no error.

Now bend a wire to the shape shown in the cut and

fasten a small piece of phosphorus to it, working under water in a tin pan or other large vessel. Transfer the wire to the jar and push the phosphorus up into the air in the graduated tube, taking care to keep the open end of the tube constantly under water. Let the apparatus stand for an hour and examine it. If the water has reached the phosphorus raise the latter so as to keep it in the air. Next morning pull the phosphorus down into the water and make the level of the water inside and outside the same. This must be done without touching the tube with the fingers. Read the volume of the residual gas. The loss in volume is *oxygen*. Calculate the percentage of oxygen by volume in the air thus:

$$\frac{\text{Volume of oxygen} \times 100}{\text{Volume of air}} = \text{per cent of oxygen.}$$

Plunge a burning splinter into the gas remaining in the tube.

In this method errors result from changes in atmospheric pressure and temperature between the two measurements. If a more accurate result is desired, a thermometer must be placed in the jar ten minutes before the volume of the air is measured in the first place. Temperature and barometer pressure must be read and recorded and the volume of the air reduced to standard conditions by the method explained in the Appendix. Temperature and pressure must again be read at the time the residual gas is measured and its volume also reduced. Then this reduced volume is subtracted from the other and the calculation made as before.

EXPERIMENT 49*b*.—**Analysis of the air. Alternative method.**[1]—Use a graduated tube holding 50 c.c., like that

[1] This method is at least as accurate as that described in 49*a*, and much more rapid. It has the advantage of not requiring the use of phosphorus.

employed in Experiment 43. Ascertain the capacity of the ungraduated portion. Have ready a cork which fits the tube *tightly*. If a wooden cork, it must be well rolled under the foot. A rubber cork without any perforation is best. Make a measuring test-tube to hold 10 c.c., as described in Experiment 43. It need not be dry.

Dissolve a small piece of potassium hydroxide in about 15 c.c. of water in a beaker and add to the liquid about 0.5 gram of *pyrogallic acid. Use the solution at once. It spoils rapidly.* Take 10 c.c. of this liquid in your measuring test-tube, pour it into the graduated tube, and instantly cork the tube tightly. Make the liquid flow from one end of the tube to the other twenty times or more. Place it in a deep cylinder of water, cork down, remove the cork, allow to stand five minutes, equalize the levels without touching the part of the tube containing the nitrogen with the hand and read the volume. Calculate thus:

Capacity of tube to end of graduations.....	50 c.c.
Capacity of ungraduated portion.....	9 "
Total capacity.....	59 "

Since 10 c.c. of liquid was introduced, the volume of air taken was 59 — 10, or 49 c.c.

Volume of gas left in the tube, 38.5 c.c.

The loss in volume is oxygen. This is

$$49 - 38.5 = 10.5 \text{ c.c.},$$

and the percentage of oxygen is

$$\frac{10.5 \times 100}{48.5} = 21.4 \text{ per cent.}$$

EXPERIMENT 50.—**Eudiometric analysis of air.**—In this experiment the U-shaped eudiometer described in Experiment 9 is used. If no Kipp generator is available, have ready a wide test-tube with a delivery tube ending in a short piece of rubber tubing. This test-tube should con-

tain a little zinc covered with water and is to be used as a source of hydrogen.

Clamp the eudiometer (Fig. 9) vertically and fill both limbs of it with mercury. Open the stop-cock *S*, place a dry, clean beaker under *C* and allow mercury to run out until about 20 c.c. of air has entered the tube. Close *C* and measure and record the volume of the air. Verify the measurement. Turn *S* so that the air in the eudiometer is isolated and so that gas passed in at *T* will escape at *E*. Pass hydrogen through the stop-cock for three minutes, allowing it to escape into the air. Then pass not more than 15 c.c. nor less than 10 c.c. of hydrogen into the eudiometer, turn the stop-cock so as to close the eudiometer, and *at once* remove the generator. Equalize the mercury levels in both limbs by running out mercury carefully from *C*. If too much mercury is run out, pour some in at the top of the open limb and again equalize. Read off the volume and record it as air + hydrogen.

Press the thumb tightly on the open end of the eudiometer and pass the spark from an induction-coil excited by three Edison-Lalande cells, two bichromate cells, or in any other suitable way. What is the cause of the explosion? Pour mercury in the open limb until the level is slightly higher than that in the other, and let the apparatus stand for five minutes. (?) Equalize the levels exactly by running out mercury from *C*, and read and record the "volume after explosion." Calculate as in the following example:

Volume of air taken.....	25	c.c.
Volume of air + hydrogen.....	39	"
Volume after explosion.....	23.4	"
The loss in volume after explosion must be	15.6	"

Now this contraction is due to the disappearance of oxygen and hydrogen to form water, and in this reaction

one volume of oxygen and two of hydrogen disappear. Hence the volume of the oxygen in the air taken must be $\frac{15.6}{3} = 5.2$ c.c., and the percentage must be $\frac{5.2 \times 100}{25} =$ 20.8 per cent.

This method of analyzing the air is employed in actual work, since it is rapid, and, with proper care, accurate.

EXPERIMENT 51.—**Absorption of the oxygen of the air by metals.**—The method is to pass air over a column of copper filings heated nearly to redness in a tube similar to that used in Experiment 10, Fig. 11. The tube should be about 30 cm. (12 inches) long, and about 20 cm. (8 inches) of it should be heated nearly to redness by a wing-top burner with several chimneys. The copper filings are held in place by loose plugs of asbestos, and after being filled the tube should be held horizontal and tapped gently on the table to make a channel along the upper portion.

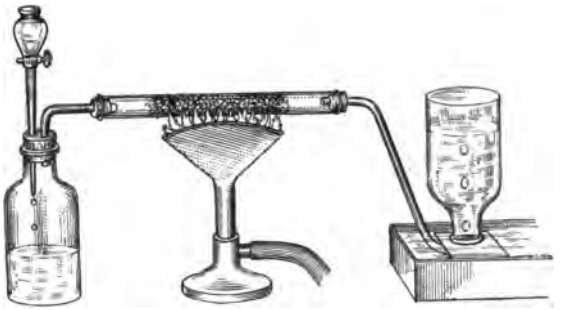

FIG. 21.

One end of the tube is connected with a bottle out of which air is expelled by water, which is allowed to drop into the bottle from a separating funnel which passes through the stopper. The air passes over the hot copper, which absorbs the oxygen, and the nitrogen is collected over water. The apparatus is shown in Fig. 21. The bottle from which the air is expelled and that in which the nitrogen is col-

lected should be of the same size and shape. The speed of the current of air depends upon the rate at which the water is allowed to fall upon the separating funnel. *It should be as slow as time permits.*

What change takes place in the copper? At which end of the tube does the change begin? About what fraction of the air disappears?

When the experiment is finished, slip a glass plate over the bottle of nitrogen and remove it from the water. Plunge a burning piece of wood or a lighted candle into the gas. (?) What is the effect of the nitrogen of the air upon combustion processes, and how is this effect caused?

Repeat the whole experiment, using iron filings instead of copper. Explain exactly what takes place.

EXPERIMENT 52.—Expose some clear lime-water in a beaker to the air for some time. Look for a change at the surface of the lime-water. What does the result prove? What are the sources of the carbon dioxide of the air and what becomes of it?

EXPERIMENT 53.—**Nitrogen.**—Put about 5 grams of ammonium chloride and the same quantity of powdered sodium nitrite, $NaNO_2$, in a small flask and add enough water to make a thin paste. Set up the apparatus as shown in Fig. 29, Part I. Apply a gentle heat and collect the nitrogen over water. *Stop heating as soon as the reaction begins or it will become too violent.* If the evolution of gas threatens to become too energetic, immerse the generating flask *for an instant only* in a pan of cold water, which should be in readiness. This will immediately quiet it. *Long immersion will cause water to flow back into the flask and spoil the experiment.*

What are the physical properties of the substance you have prepared? Does it burn? Does it support the combustion of a candle?

Questions

1. What constituents of the air are essential to life, and why?

2. What are the reasons for the belief that the air is a mixture, not a compound?

3. Describe an experiment which proves that the air contains the vapor of water.

4. Nitrogen is lighter than oxygen and water-vapor only about half as dense. On the other hand, carbon dioxide is much denser than oxygen. Why is it that all these substances remain uniformly mixed in the air and do not separate into layers according to density?

5. What is the effect of the nitrogen of the air upon combustion processes, and why?

6. Is there any difference between nitrogen obtained from the atmosphere according to Experiment 51 and that prepared by chemical methods as in Experiment 53? Why?

7. What elements are most abundant in the bodies of animals and plants?

Problems

43. 16.7 c.c. of air were confined over mercury in a eudiometer and enough hydrogen added to make the volume 30 c.c. After explosion, the volume was 19.5 c.c. What percentage of oxygen by volume did the air contain?

44. 20 c.c. of air are mixed with 10 c.c. of hydrogen and the spark is passed. After the explosion, what volumes of what gases remain in the tube?

In solving this problem assume that the air contains 21 per cent by volume of oxygen.

45. 2 liters of air were passed over hot copper. The increase in weight of the copper was .6 gram. What was the percentage of oxygen by weight in the air?

Assume the weight of 1 liter of air to be 1.293 grams.

46. What weight of nitrogen can be obtained by heating 13.8 grams of sodium nitrate, $NaNO_3$, with the required quantity of ammonium chloride, NH_4Cl? How much ammonium chloride is needed? What weights of salt and of water are formed?

CHAPTER XIV

AMMONIA

EXPERIMENT 54.—**Preparation of ammonia gas.**—Place about 150 c.c. of ammonium hydroxide in a 500 c.c. flask. Arrange the apparatus as shown in Fig. 31, Part I. A little mercury should be placed in the bend of the safety-tube. A piece of asbestos board must be placed under the flask. Use glass tubing as far as possible in constructing the apparatus. A rubber cork can be used in the flask.

Apply a gentle heat and collect the ammonia by *upward* displacement in bottles, which must be *absolutely dry.* Why? The tube must run up to the bottom of the inverted bottle. From time to time hold a drop of hydrochloric acid on a glass rod to the mouth of the bottle. Dense white fumes indicate that the latter is full and ammonia is escaping. Collect three bottles of the gas and stand them on glass plates in inverted position. Hold the exit tube of your apparatus to the burner flame. How does the gas behave as regards combustion? Does it produce a continuous flame? Can it be made to do so? What are the products of its combustion?

Plunge the mouth of a bottle filled with the gas under water and move it so as to bring the water in contact with the gas. (?) *Remove the glass plate before placing the mouth of the bottle under the water.*

Slowly introduce a lighted candle held on a wire into an inverted bottle of the gas, as you did with hydrogen. Does it support combustion? How does it behave on the instant of contact with the candle-flame?

Fill a dry bottle with hydrochloric-acid gas by downward displacement. The gas can be made by gently heating a little strong hydrochloric acid placed in a test-tube with a cork bearing a delivery tube bent downward. Bring

the bottle mouth to mouth with a bottle of ammonia gas. Let the bottles stand until the product has settled on the sides. Scrape some of it out and heat it carefully in a porcelain crucible or on a piece of platinum foil. (?) What is it?

EXPERIMENT 55.—**Ammonium salts.**—Dilute some ammonium hydroxide (ammonia water) with about three times its volume of water in a beaker. How does the liquid affect red and blue litmus paper? Does it act upon the litmus in the same way as caustic soda?

Dilute about 5 c.c. of hydrochloric acid with three times its volume of water in a dish and investigate its behavior with both kinds of litmus paper. *All acids affect litmus paper in the same way as hydrochloric acid.* Ammonium hydroxide and sodium hydroxide are *bases. All soluble bases affect litmus paper in the same way.*

Now add your dilute ammonium hydroxide, drop by drop, to the hydrochloric acid in the dish, stirring constantly. From time to time dip small pieces of blue litmus paper into the liquid. When the latter is no longer turned red you will find that a piece of red litmus paper is turned blue. At the same time the liquid acquires a faint odor of ammonia. Why?

Evaporate the liquid in the dish to dryness. Examine the residue. Is it ammonium chloride? Heat some of it on a piece of platinum foil or in a crucible. Write the equation.

Repeat the whole experiment, using nitric acid instead of hydrochloric. Write the equation.

Dissolve some ammonium chloride in water in a test-tube. Add a fragment of sodium hydroxide to the liquid and heat gently. Notice the odor of the gas given off. Hold a fragment of red litmus paper in the gas. Repeat, using potassium hydroxide instead of sodium hydroxide. In the same way investigate the behavior of ammonium

nitrate with sodium hydroxide and potassium hydroxide. Notice that *strong bases liberate ammonia from ammonium salts.* Write equations for the four chemical changes which have taken place.

Study the effect of heat upon small quantities of ammonium nitrate, chloride, and sulphate. Use a porcelain crucible or a piece of platinum foil.

EXPERIMENT 56.—Dilute about 30 c.c. of ammonia with about ten times its volume of water, in a flask. Shake up the liquid with salt until no more dissolves. Fill the apparatus employed in the electrolysis of water (Fig. 6) with the clear liquid and pass the current. What two gases collect? Try to burn each. Close the stop-cocks for a time and ascertain the relative quantities of the gases. If the apparatus is not graduated, measure the length of the columns of gas with a meter scale.

In this experiment the salt is added only to assist in conducting the current through the liquid, and thus make the process more rapid.

QUESTIONS

1. What does the presence of ammonia indicate with regard to the fitness of a sample of water for drinking?

2. Explain the fact that the air of stables often smells strongly of ammonia.

3. Remembering that the density of nitrogen referred to hydrogen is 14, and that the symbol N means 14 parts of nitrogen by weight, deduce from Experiment 56 the formula of ammonia.

4. Devise two methods of distinguishing an ammonium salt from a sodium or a potassium salt.

5. How would you convert the hydrogen of ammonia into water?

6. How would you convert the hydrogen of water into ammonia? (See paragraph 139, Part I.)

7. How would you convert free nitrogen into ammonium nitrate? (See p. 105, Part I.)

8. What are some of the uses of ammonia water? *Liquid ammonia* is largely employed for one purpose only. What is it?

9. How could you make ammonia, taking the hydrogen from sulphuric acid and the nitrogen from ammonium nitrite, NH_4NO_2?

10. Explain exactly what is meant by a radical.

Problems

47. 856 grams of ammonium chloride are heated with sodium hydroxide. How much ammonia by weight escapes?

48. Calculate the percentage composition of

a. Ammonium chloride, NH_4Cl.
b. Ammonium nitrate, NH_4NO_3.
c. Ammonium sulphate, $(NH_4)_2SO_4$.
d. Ammonium hydroxide, NH_4OH.

49. When a stream of electric sparks is passed through ammonia it is decomposed, two volumes yielding one volume of nitrogen and three volumes of hydrogen. What volumes of nitrogen and hydrogen are formed when 300 c.c. of ammonia are treated in this way?

50. 100 c.c. of ammonia are decomposed by a stream of sparks. (*a*) What volume of oxygen must be added to the resulting mixture to combine with the hydrogen and produce water? (*b*) After the water has condensed, what gas will remain in the tube, and how much?

51. 32 c.c. of ammonia are decomposed by sparks, 50 c.c. of oxygen are added, and the mixture is caused to explode. What volumes of what gases are left?

52. An unknown volume of ammonia is decomposed in a eudiometer, an unknown volume of oxygen is mixed with it, and the mixture exploded. After the explosion the contraction in volume is 18 c.c. and the tube still contains some oxygen. (*a*) What volume of ammonia was taken in the first place, and (*b*) what volume of nitrogen was left in the tube?

53. Ammonium chloride is heated in a flask with sodium hydroxide and the ammonia passed into 31.5 grams of pure nitric acid. How much ammonium chloride must be used in order to convert all the nitric acid into ammonium nitrate?

CHAPTER XV

COMPOUNDS OF NITROGEN AND OXYGEN

EXPERIMENT 57.—**Nitrous oxide.**—Fill a wide test-tube one-third with ammonium nitrate. Clamp it in an inclined position, and insert a perforated cork with a delivery tube. It is advisable to pass the gas through an empty wide test-tube closed by a doubly perforated cork before collecting it. Apply a gentle heat and collect the gas over warm water. *The evolution of gas must be slow —a bubble or two a second. If it becomes too rapid explosions result.* This can be easily controlled by lowering or removing the flame. Stop heating and remove the cork before the ammonium nitrate is exhausted. *Explosions sometimes occur when the quantity of substance becomes small.* If the ammonium nitrate shows signs of giving out before you have enough gas, disconnect, add more of it directly to the liquid in the test-tube, and resume heating cautiously. Is there evidence of the formation of any product besides the gas? What?

Collect three bottles and two test-tubes full of the gas. Use one bottle to determine color, odor, and taste. (?) Try the spark test in a test-tube. Ascertain if the gas is soluble in water. Plunge a lighted candle into a bottle of it. (?) Set fire to some sulphur in a deflagrating spoon, and *at the instant the sulphur begins to burn* plunge it into the third bottle of nitrous oxide. It should be extinguished. Cover the bottle, and heat the sulphur until it burns vigorously. Plunge it again into the gas. (?) Is nitrous oxide easy or difficult to decompose into its elements? Which has probably the higher temperature, the candle-flame or the flame of burning sulphur?

EXPERIMENT 58.—**Nitric oxide.**—*Make no attempt to ascertain the odor of nitric oxide. It must not be inhaled,*

and experiments with it should be carried out under the hood, if possible. If you do not work under the hood, keep the bottles of the gas covered during combustions and let the products escape out of the windows.

Nitric oxide is made in the same apparatus which is used for generating hydrogen. Fill the generator to the depth of about 1 cm. with copper clippings or cut pieces of sheet copper. In another vessel, dilute some nitric acid with twice its volume of water. Cool this liquid and pour it upon the copper until the generator is one-third filled. Collect the gas over water. It is well to stand the generator in cold water. If the evolution of gas becomes too energetic, pour a little water down the funnel-tube. Collect four bottles and one test-tube of the gas. Then stop, as even the first gas which comes off is not quite pure and the later portions contain large quantities of *nitrous* oxide and other impurities. Leave the bottles standing in the water.

Take the apparatus to the hood, disconnect it (*do not inhale the gas*) and pour a little of the blue liquid into a clean dish. Put a small flame under it and let it evaporate almost to dryness under the hood. What is the product? Wash off the copper in the generator and return what is left of it to the stock-bottle.

Expose a bottle of nitric oxide to the air. (?) Into another bottle slowly admit oxygen from a cylinder or gas holder. Stop after adding a few bubbles of oxygen, and shake the bottle, keeping its mouth under water; then continue adding oxygen. Explain these results.

Slip aside the cover of a third bottle and plunge to the *bottom* of the bottle some sulphur, burning vigorously in a spoon. Remove the sulphur at once, keeping the bottle covered, and do the same thing with a lighted candle. Why should the candle be placed in the bottom of the bottle? Interpret the results.

Into the fourth bottle introduce a piece of burning magnesium held in forceps. (?) Which flame has the higher temperature, that of a candle or magnesium? Which gas requires the higher temperature to decompose it, nitrous or nitric oxide? Construct a tabular view of the properties of the two gases.

EXPERIMENT 59.—**Nitrogen peroxide.**—Place in a dry sealed tube of hard glass lead nitrate to the depth of about 1 cm. and heat gently. The crackling which often occurs is due to the fact that the crystals inclose small drops of water and fly to pieces when this water is vaporized by heat. What two gases are given off? What is left in the tube? Bring into the mouth of the tube a piece of moist blue litmus paper. (?) The result is due to the fact that nitrogen peroxide reacts with water, producing *nitric acid.*

EXPERIMENT 60.—**Formation of nitric acid from water and air under the influence of electric discharges.**—Select a small glass flask or bottle and a 2-hole rubber stopper which fits it. The holes in the stopper should be small. Pass through each hole a platinum wire and jam a piece of glass rod into each hole to fix the position of the wires. Both wires should reach nearly to the bottom of the bottle, and should be bent toward each other so that the ends are less than 1 cm. apart. They must not touch. The bottom of the bottle must be wet. Throw a piece of moist blue litmus paper into the bottle, insert tightly the cork with the wires, and pass a stream of sparks from a coil for half an hour, or until the litmus paper changes color. The result is due to the production of *nitric acid.* What light does this experiment throw upon the presence of nitric acid in rain (paragraph 132, Part I)? Upon the supply of nitrogen to plants?

EXPERIMENT 61.—**Preparation of nitric acid.**—*The gases given off when nitric acid acts upon metals and the vapor of nitric acid itself are poisonous and must*

not be inhaled. Do not get nitric acid upon the skin or clothing.

Fit up the apparatus shown in Fig. 22. Roughly weigh on the platform scales enough potassium nitrate to fill the retort about one-third. Weigh off in a beaker an equal quantity of strong sulphuric acid. Introduce the potassium nitrate into the retort by means of a piece of paper. Add the sulphuric acid, insert the *glass stopper* (not a rubber or wooden cork, which would be rapidly destroyed by the nitric acid) and heat gently. Only vapor, not solid or liquid, must pass over. Why?

FIG. 22.

Examine the nitric acid which collects and record its properties. Use it in the following experiments. Cover a piece of zinc with water in a test-tube and slowly add nitric acid. Do the same thing with a little iron filings and a piece of magnesium. In these reactions water is produced, together with the nitrate of the metal used. The gas produced may be nitrous oxide, nitric oxide, or nitrogen peroxide according to the temperature and strength of the acid. Do not try to write the equations, which are quite difficult. Simply remember the general character of the action and the way in which it differs from the action of sulphuric or hydrochloric acid on the same metals.

Drop a fragment of tin into strong nitric acid in a test-tube. What are the products (paragraph 154, Part I)?

Drop a piece of cork into nitric acid and let it remain

there for a time. Let a drop of nitric acid fall upon a colored fabric of any kind. What is the action of nitric acid upon organic matter? Would the formula of the acid lead you to expect such action?

EXPERIMENT 62.—**Aqua regia.**—Pick up a small piece of gold leaf with the end of a wet glass rod, and rinse it into a test-tube with water. In the same way place another piece of gold-leaf in another test-tube. Add to one tube nitric acid and to the other hydrochloric acid. Heat both tubes. Does the gold dissolve. Pour the contents of either tube into the other, and go on heating. (?) The product is gold chloride, and the gold dissolves because the oxygen of the nitric acid liberates chlorine from the hydrochloric acid and the chlorine attacks the gold. Suggest another substance which could be used with the hydrochloric acid in place of the nitric acid, though not so conveniently. The mixture of nitric and hydrochloric acids will also dissolve platinum.

EXPERIMENT 63.—**The nitrates.**—Examine potassium nitrate and sodium nitrate. Try the solubility of each in water. Place a drop of each solution on a clean glass plate, let it evaporate to dryness, and examine the crystals left with a good lens. Record your observations.

In a small iron or nickel dish melt enough potassium nitrate to half fill it and heat the melted salt with the full power of the burner flame. Be careful not to upset the dish, as the melted substance would burn deeply into the table. Throw into it a piece of charcoal which has been heated at one corner to redness. Stand aside as the melted substance may sputter. Throw a fragment of sulphur into the dish. Is the oxygen of potassium nitrate firmly or loosely held? Does the result throw any light upon the behavior of the potassium nitrate in gunpowder? Weigh out 6 grams of powdered potassium nitrate, 1 gram of powdered charcoal, and 1 gram of flowers of sulphur.

Mix thoroughly by pouring from one paper to another—NOT IN A MORTAR. Place the heap upon a piece of wood, thrust a piece of filter-paper into it, light the paper and stand aside. If the paper fails to ignite the mixture, wait until the last spark of the burning paper is extinguished before approaching the heap to make a second attempt. Explain.

Place a crystal of silver nitrate on charcoal and let the burner flame play upon it. (?)

EXPERIMENT 64.—**The nitrites.**—Examine sodium nitrite. Dissolve some of it in water. Add to a portion of the solution a drop or two of silver nitrate solution. To another portion add carefully a little sulphuric acid. Interpret the result after reading paragraph 159, p. 117, Part I.

QUESTIONS

1. What is the explanation of the peculiar conduct of nitrous and nitric oxides toward combustible substances?

2. Vitali, an Italian experimenter, plunged cats into nitric oxide, and observed that they died in convulsions. Why is it that this experiment does *not* furnish any information about the action of nitric oxide upon the animal body?

3. Sodium nitrate is much cheaper than the potassium salt; but supposing that the two sold at the same price, why would it be more profitable to use the sodium compound in making nitric acid?

4. Why do railroads object to transporting nitric acid if packed in glass vessels?

5. Describe in general terms the action of nitric acid upon metals, and explain what becomes of the hydrogen of the acid.

6. Nitrous oxide yields the spark test and supports combustion brilliantly. How could it be distinguished from oxygen?

7. Nitrogen and chlorine do not combine directly to produce nitrogen chloride; but supposing they did, would heat or cold be produced? Why?

8. Why is it impossible to employ nitrogen chloride practically as an explosive?

Problems

54. What weight of nitric acid containing 80 per cent HNO_3 is necessary to dissolve 10 grams of cupric oxide?

$$CuO + 2HNO_3 = Cu(NO_3)_2 + H_2O.$$

55. What weight of pure nitric acid would contain 50 grams of oxygen?

56. Assuming that the density of pure nitric acid is 1.5, how much oxygen do 3 liters of it contain?

57. How much nitric acid can be obtained (*a*) by heating 200 kilos of sodium nitrate with sulphuric acid; (*b*) by heating 200 kilos of potassium nitrate with sulphuric acid?

58. I require 120 grams of cupric oxide. How much crystallized cupric nitrate must be heated to redness to make it?

$$Cu(NO_3)_2 3H_2O = CuO + 2NO_2 + O + 3H_2O.$$

59. How much nitric acid is needed to convert 400 grams of potassium hydroxide into potassium nitrate?

$$KOH + HNO_3 = KNO_3 + H_2O.$$

60. Calculate the percentage composition of N_2O_5.

61. Calculate the formula of a compound having the following composition:

9.09 per cent nitrogen;
20.77 " oxygen;
70.13 " silver.

What is the name of the compound and how could you make it?

CHAPTER XVI

ATOMIC AND MOLECULAR WEIGHTS—AVOGADRO'S LAW

No experiments.

Questions

Use the utmost precision in answering these questions. The subject is very important.

1. Explain exactly what is meant by the term *molecular weight.*
2. What is the exact difference of meaning between the follow-

ing two formulas—*a*. NO_2; *b*. N_2O_4? Answer this question first in the language of grams and liters. Then answer again in the language of the atomic theory.

3. Why is the density of a gas or vapor referred to hydrogen equal to one half the molecular weight? Answer this question first in the language of grams and liters; then in the language of the atomic theory.

4. Explain how to calculate from the formula the weight of one liter of any gas or vapor.

5. Explain how to calculate from the formula the volume of one gram of any gas or vapor. For what temperature and pressure are the results obtained in 4 and 5 good?

6. Show that 18 grams of water, H_2O,
58.5 " " salt, NaCl,
63 " " nitric acid, HNO_3,
2 " " hydrogen, H_2,
must all contain the same number of molecules.

7. Explain Avogadro's Law.

8. What is the law of simple volume ratios? Why would the atomic theory lead us to expect it to be true?

9. What basis of fact is required in order to double a formula?

Problems

62. Calculate the molecular weights of the following compounds:

a. Sugar, $C_{12}H_{22}O_{11}$;
b. Bismuth nitrate, $Bi(NO_3)_3$;
c. Nitroglycerin, $C_3H_5(NO_3)_3$;
d. Glucose, $C_6H_{12}O_6$.

63. What is the weight of 28 liters (*a*) of nitrous oxide? (*b*) of nitric oxide?

64. What is the volume (*a*) of 11 grams of nitrous oxide? (*b*) of 5 grams of nitric oxide?

65. What volume of hydrogen is needed to form water with the hydrogen (*a*) of 22 liters of nitrous oxide? (*b*) of 22 grams of nitrous oxide?

$$N_2O + H_2 = H_2O + N_2.$$

66. What is the volume of 13 grams of nitric oxide?

67. What volume of nitrous oxide—measured at standard con-

ditions—can be made from 80 grams of ammonium nitrate? Solve by inspection.

68. How much ammonium nitrate is needed to make 80 liters of nitrous oxide?

69. How much ammonium nitrate is needed to make 4,000 c.c. of nitrous oxide?

70. How much copper is needed to produce 80 liters of nitric oxide?

71. How much ammonium nitrate is needed to produce 10 liters of nitrous oxide?

72. What volume of nitrous oxide at 15° C. and 700 mm. can be obtained by heating 200 grams of ammonium nitrate?

73. What volume of nitric oxide at 13° and 740 mm. is obtained by dissolving 80 grams of copper in nitric acid?

CHAPTER XVII

ACIDS, BASES, AND SALTS—METALS AND NON-METALS

EXPERIMENT 65.—**The reaction of sodium hydroxide with hydrochloric acid.**—Dissolve a piece of sodium hydroxide about 5 cm. (2 inches) long in 10 c.c. of water in a beaker. Carefully and slowly add strong hydrochloric acid, stirring constantly. Is there evidence of energetic action? A white solid separates. What is it? What else was formed at the same time?

Let the solid settle and pour off the liquid. Dissolve the solid in the smallest possible quantity of water, transfer it to a dish, and evaporate it slowly to dryness. When the residue is perfectly dry, let it cool and examine it. Taste it. (?) Write the equation.

EXPERIMENT 66.—**The reaction of sodium hydroxide with sulphuric acid.**—Make a *little* very strong solution of sodium hydroxide in a test-tube. Place about 2 c.c. of the liquid in another test-tube and clamp it vertically. Drop strong sulphuric acid into it—*one drop at a time*—

from a tube drawn out to a jet. Be careful. The reaction is violent. What are the two products? Is heat evolved or absorbed? Write the equation.

An acid and a base always react, forming a salt and water when they are brought together, but the reaction is not always as energetic as in this instance.

EXPERIMENT 67.—**Properties of acids and bases.**—Prepare very dilute hydrochloric, nitric, sulphuric, and acetic acids by diluting the acids with about 100 times their volume of water. Make at least 300 c.c. of the hydrochloric and nitric: a small quantity only of the others. Taste each of these liquids. Put a drop on red litmus paper. On blue litmus paper. On turmeric paper.

Dilute some ammonia with about 100 volumes of water. Dissolve about 3 grams of sodium hydroxide in 300 c.c. of water. Dissolve 3 grams of potassium hydroxide in 300 c.c. of water. Obtain some clear lime-water—this is already sufficiently dilute. Taste these liquids, rub them between the fingers, and test them with the same papers which were employed for the acids.

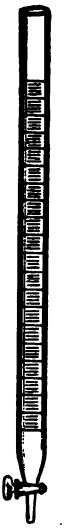

FIG. 23.

EXPERIMENT 68.—**Neutralization.**—(1) Clamp two *burettes* (Fig. 23) vertically side by side. Fill one with dilute hydrochloric acid (from Experiment 67), and let the liquid slowly drop out until the bottom of the meniscus (Fig. 7, Part II) is at *O*. Fill the other to *O* with sodium hydroxide solution (from Experiment 67). Measure out 10 c.c. of hydrochloric acid into a clean beaker, and add a few drops of litmus solution—enough to color the liquid *faintly*. Now allow sodium hydroxide solution to run into the liquid, stirring with a glass rod until a change in color occurs. This liquid must fall one drop at a time, and the change in color should be produced by a single drop.

Before this last drop was added the liquid was *acid*—that is, it contained $\overset{+}{H}$ ions; afterward it was *alkaline,* it contained $\overline{OH}$ ions. Read off the level of the liquid in the burette, and calculate how much of your sodium hydroxide would be required for 1 c.c. of your acid.

Now take 20 c.c. of the hydrochloric acid and add, instead of litmus, 2 or 3 drops of a solution of phenol phthalein. The liquid will remain colorless. Stir the liquid constantly and find out just how much of the sodium hydroxide solution is required to make it alkaline. In this case the $\overline{OH}$ ions, when there are any present, will produce a red color.

Take the reading and calculate again how much of the sodium hydroxide is required for 1 c.c. of the acid. Do the two results agree? Repeat until you are sure that there is a fixed relation between the quantities of the two liquids which react.

What becomes of the $\overline{OH}$ ions which are added before the color change occurs?

(2) Repeat, using potassium hydroxide and nitric acid —the dilute liquids from Experiment 67. This time measure off 10 c.c. of the base and add the acid drop by drop from its burette. Use litmus as the indicator. Take 20 c.c. of the base the second time and use a few drops of a solution of *cochineal,* easily made by digesting the crushed insects in alcohol for some hours. It is orange when acid and violet when alkaline.

If time permits, repeat (1), using potassium hydroxide instead of sodium hydroxide with the hydrochloric acid. Calculate the volume of your potassium hydroxide which would be equivalent in neutralizing power to 1 c.c. of your sodium hydroxide. Thus suppose that

8 c.c. NaOH are equivalent to 10 c.c. HCl,
and 12 " KOH " " " 10 " "
clearly 8 " NaOH " " " 12 " KOH,
and 1 " NaOH is " " $\frac{12}{8}$, or 1.5 c.c. KOH.

Repeat (2), using sodium hydroxide instead of potassium hydroxide with the nitric acid. Calculate again the volume of your KOH solution, which is equivalent to 1 c.c. of your NaOH. Does the result agree with that obtained above? If not, repeat. Since the burette only reads to 0.1 c.c., do not look for absolute agreement.

EXPERIMENT 69. — **Electrolysis.** — Place some dilute solution of copper sulphate, $CuSO_4$, in a beaker. In the liquid place two pieces of platinum foil each about 2 cm. square. Each is attached to an insulated copper wire which leads to the terminal of an Edison-Lalande cell, or some other source of the electric current. It is well to have a cheap galvanometer in the circuit. Pass the current for fifteen minutes and observe the result. The current evidently passes through the liquid. What carries it? Does the copper go to the negative or positive pole? Why? Reverse the direction of the current for a time. (?) The platinum foil can be cleaned with a little nitric acid.

Dissolve some sugar in about twenty times its weight of distilled water, and introduce the electrodes. A galvanometer must be in the circuit. Does the sugar solution conduct? Sugar contains much hydrogen. Its formula is $C_{12}H_{22}O_{11}$. Is it an acid? Does it yield $\overset{+}{H}$ ions when dissolved in water? Verify your conclusion by testing the solution with red and blue litmus paper.

Dilute some alcohol with about ten times its volume of water and introduce the electrodes. Does it conduct? Alcohol is the hydroxide of a radical. Its formula is C_2H_5OH. Is it a base? Does it dissociate, yielding $O\overset{-}{H}$

ions when mixed with water? Verify your conclusion by testing the liquid with both kinds of litmus paper.

Add very dilute sulphuric acid (about 1 to 100 parts of water) to some blue litmus solution, one drop at a time, stirring constantly until the color is purple, halfway between blue and red. If you put in too much acid, making the color red, add very dilute ammonia. Dissolve some sodium sulphate in water, color the liquid with this neutral litmus, place the liquid in a U-tube, and pass the current for a time. Is any metal deposited at the negative pole? What is formed there? Salts of the metals of the sodium and calcium groups always behave in this way when the current is passed through water solutions. What happens at the positive pole?

EXPERIMENT 70.—**Metals and non-metals.**—As a typical solid non-metal, examine *sulphur*. As a typical metal, examine *lead*. Notice the difference in luster. If the lead is tarnished, scrape it to expose a fresh surface. Rub the sulphur vigorously on the coat-sleeve, and bring it near to some fragments of paper. Notice that it receives a charge and retains it. This shows that the sulphur is a very bad conductor of electricity. Treat the lead in a similar way, and notice that the charge is immediately conducted away through the lead and the body to the earth. The lead *conducts* the current. The same difference exists between the capacity of the two substances for conducting *heat*.

Investigate the tenacity of the two by endeavoring to pull apart a fragment with the hands. (?) Investigate their behavior under the hammer. What are the chief physical differences between metals and non-metals? What are the chemical differences?

As an example of an element on the border-line between the two classes examine *antimony*. Has it a metallic luster? Is it malleable?

Questions

1. What is an acid? What is a base? What happens when the two are brought together? What is a salt?

2. How do acids act upon oxides? Upon carbonates? Why is it that no gas escapes when an acid acts upon an oxide or a hydroxide?

3. What is an *ion*? Give two examples of ions which are single atoms. Give two examples of ions which are groups of atoms.

4. What is the exact difference between chlorine ions and chlorine gas?

5. Sodium, like other metals, contains but a single atom in its molecule. What, then, is the difference between metallic sodium and sodium ions?

6. According to the idea of electrolytic dissociation, when the substances are all dissolved in water,

Sodium hydroxide consists of $\overset{+}{\mathrm{Na}}$ and $\overset{-}{\mathrm{OH}}$;

Hydrochloric acid " " $\overset{+}{\mathrm{H}}$ and $\overset{-}{\mathrm{Cl}}$;

Sodium chloride " " $\overset{+}{\mathrm{Na}}$ and $\overset{-}{\mathrm{Cl}}$.

Between what two ions does the actual reaction occur when NaOH and HCl are brought together? Discuss exactly what happens to all four ions.

7. Using the dissociation-idea, explain the fact that all salts of the same metal when dissolved in abundant water produce the same color.

8. Is electrolytic dissociation a fact or a supposition?

Problems

74. 10 grams of pure sodium hydroxide are dissolved in water. *a*. How much nitric acid must be added to make the solution neutral? *b*. How much sodium nitrate would be obtained if this was done?

75. I have a solution which contains just 40 grams of pure sodium hydroxide in 1 liter. Calculate the quantities by weight of (*a*) HCl, (*b*) HNO_3, and (*c*) H_2SO_4, which will be required to neutralize 1 c.c. of it.

76. 15.75 grams of nitric acid are mixed with 23.25 grams of

sodium hydroxide, both dissolved in water. What two compounds does the solution contain and how much?

77. In ascertaining the strength of a dilute solution of HCl, 50 c.c. of it were measured out and neutralized with a solution of sodium hydroxide containing .003 gram of NaOH in 1 c.c. 40 c.c. of the sodium hydroxide solution was required. What weight of HCl was contained in 1 c.c. of the hydrochloric acid?

78. 30 c.c. of a solution of potassium hydroxide containing .01 gram of KOH in 1 c.c. was needed to neutralize 40 c.c. of a solution of HCl. How much HCl did 15 c.c. of the hydrochloric-acid solution contain?

79. 20 c.c. of a solution containing .005 gram of KOH in 1 c.c. just neutralized 20 c.c. of a solution of hydrochloric acid. How much HCl did 15 c.c. of the latter contain?

CHAPTER XVIII

THE SODIUM GROUP

EXPERIMENT 71.—**Potassium.**—*Potassium is preserved under naphtha, and the bottle containing it must be kept stoppered and not opened in the vicinity of a flame. It catches fire on contact with water, and all apparatus used in handling it, as well as the desk, must be dry. It must not be touched with the fingers. When exposed to the air it takes fire spontaneously after a time; therefore it must not be put away under the desk or left lying around. When thrown into water it reacts explosively, and a glass plate should be used to protect the eyes when the experiment is performed.*

Examine a piece of potassium. Cut it and notice the luster. Is it permanent? Why? Add 2 or 3 drops of a solution of phenol phthalein to a bottle half full of water, and throw in a piece of potassium half the size of a small pea (*no larger*). What happens? What sub-

stances are formed? Does the liquid contain hydroxyl ions?

Scrape a hollow in a fragment of charcoal, place in it a fragment of potassium half the size of a pea, and let the burner flame play upon it. Notice the flame color.

Make in the form of a small table a comparison between potassium and sodium. Include in this comparison luster, permanence of luster, hardness, flame color, behavior with water, and any other properties that may appear important.

EXPERIMENT 72.—**Potassium compounds.**—Examine potassium hydroxide, chloride, chlorate, bromide, and carbonate. Test each for solubility in water. Make the flame test with each, using a clean iron wire, as with sodium. Clean the wire after each test by dipping it in a little hydrochloric acid in a beaker and holding it in the flame.

Make a small quantity of a mixture of potassium and sodium chlorides and apply the flame test to the mixture. Look at the flame of the mixture through blue glass.

Make the flame test with lithium chloride, LiCl. If a spectroscope is available, use it to observe the spectra of potassium, sodium, and lithium chlorides separately.

QUESTIONS

1. What is the connection of potassium with plant life?

2. From the standpoint of electrolytic dissociation, what really happens when potassium acts upon water is that the potassium passes from one condition to another. Explain this statement.

3. From the same standpoint, explain why, if a metal acts violently upon water, we should expect its hydroxide to be a strong base and its salts to be good conductors of the electric current.

PROBLEMS

80. How much potassium is required to liberate from water enough hydrogen to combine with 3 grams of oxygen?

81. What volume of oxygen is needed to combine with the hydrogen given off by the action of 9.75 grams of potassium on water?

82. How much (*a*) potassium carbonate must be heated with how much (*b*) pure charcoal to produce 9.75 grams of potassium, and (*c*) what volume of carbon monoxide would be liberated?

$$K_2CO_3 + 2C = 3CO + 2K.$$

CHAPTER XIX

THE COPPER GROUP

EXPERIMENT 73.—**Copper.**—Examine copper filings and sheet copper. Record its properties. Heat a piece of sheet copper in the Bunsen flame. Place a little copper filings in a test-tube, cover with water, and add hydrochloric acid. Is there any result? Repeat, using water and sulphuric acid. Repeat, using water and nitric acid. What is the gas given off when nitric acid dissolves copper, and why does it turn red on entering the air?

Make a spiral, about 15 cm. long, of copper wire by winding it around a glass tube. Have ready a wide test-tube and a cork to fit it. The test-tube should contain about 1 c.c. of alcohol. Hold the spiral in forceps and pass it through the burner flame until it turns black. (?) Remove it from the flame, place in the test-tube, and loosely cork the latter. When cool, remove it from the tube and examine it. It now has the true color of copper. Explain what has happened.

EXPERIMENT 74.—**Cupric oxide.**—Examine cupric oxide. Mix 3 grams of it with 0.5 gram of powdered charcoal and heat the mixture in a hard glass test-tube with a delivery tube so arranged that any gas given off must pass through lime-water in a wide test-tube. Heat gently at first and then more intensely. When the gas has been

bubbling through the lime-water for a time, introduce a lighted match into the upper part of the wide test-tube. What gas is given off? What must have happened to the copper oxide?

Cover the hard glass tube with a layer of soot from the luminous flame and let it cool completely. Examine the contents with a lens.

How does carbon affect the oxides of most metals? What is the importance of the reaction?

EXPERIMENT 75.—**The effect of cupric oxide upon organic compounds.**—Heat to redness in a hard glass test-tube, 5 grams of a mixture of cupric oxide with $\frac{1}{20}$ of its weight of sugar, both finely powdered and carefully mixed. Conduct the gases given off into the bottom of an empty test-tube which stands in a small beaker containing a freezing mixture. Use a mixture of crushed ice and salt, or else crystallized sodium sulphate with twice its weight of hydrochloric acid. This test-tube is closed with a doubly perforated cork, and from it the gases are led into a second test-tube containing clear lime-water. This tube need not be cooled and must not be corked. What is the result? What does the experiment teach about the composition of sugar? What becomes of the copper oxide? Cover the tube with soot and disconnect before cooling.

EXPERIMENT 76.—**Cupric sulphate. Cupric ions.**—Examine cupric sulphate. Heat a little of it carefully in a porcelain dish. (?) Make a solution of cupric sulphate. What is the color of cupric ions? Place a portion of the solution in a test-tube, add a piece of zinc and a few drops of sulphuric acid to hasten the action. Let stand for some time. The solid product is *copper,* finely divided and lusterless. Why does the liquid lose its color? Which has the strongest tendency to exist as ions, copper or zinc?

Repeat the experiment, using iron (a nail) in place of the zinc.

Place a small piece of clean sheet copper in a solution of silver nitrate. What is the product? Which tends most strongly to exist as ion, copper or silver? What ought, therefore, to be the effect of zinc upon silver nitrate solution? Try it with a fresh portion.

Add a drop or two of your copper sulphate solution to a test-tube half filled with water, and investigate the action of ammonia water on the liquid. This is a delicate test for cupric ions.

QUESTIONS

1. How would you convert cupric ions into metallic copper? Metallic copper into cupric ions?

2. What is the exact difference between cuprous and cupric ions?

3. The same electric current is passed in succession through a solution containing cuprous ions and one containing cupric ions. In one hour .3175 gram of copper deposits in the cupric solution. How much is deposited in the cuprous solution in the same time, and why?

PROBLEM

83. 2 grams of finely divided copper were heated in oxygen. 2.5063 grams of cupric oxide were produced. Calculate the atomic weight of copper.

CHAPTER XX

SILVER

EXPERIMENT 77.—**Silver.**—Dissolve about 0.5 gram of silver nitrate in 50 c.c. of distilled water in a beaker. Try the action of a clean piece of sheet copper upon a portion of the liquid in a test-tube. Add a drop of mercury to another portion. Let both tubes stand some time. The visible product is *silver*. What else must have been produced in both cases?

To the rest of the silver nitrate solution add a solution of sodium chloride. The precipitate is silver chloride, AgCl. Write the equation. In what previous experiment did you make and study silver chloride? What liquid dissolves it? What is the effect of light upon it? Answer these questions from your notes. Stir up the silver chloride with water, allow it to settle, and pour off the water without losing any of the silver chloride. Repeat three or four times. This is called *washing by decantation.* Place a piece of zinc in contact with the silver chloride, and add a drop or two of sulphuric acid. Allow to stand. The product is *silver.* Wash it by decantation, examine it, and record its properties.

EXPERIMENT 78.—**Photography.**—If a negative is available, examine it. What is the material of which the image consists? Describe briefly, from the beginning, the processes through which the negative has passed. Why are the lights and shadows reversed in it? If practicable, expose a piece of sensitive paper back of the negative in a printing-frame to sunlight. The paper must be inserted by dim light, and the sensitive side of the paper must be in contact with the side of the negative bearing the film. Once a minute open *half* the printing-frame in a dim light and examine the progress of the printing. When the print is finished, remove and examine it. How can the unpleasant reddish color be altered? What is *toning,* and how does it affect the chemical composition of the image? Is your picture permanent? Expose it to strong light and find out. If not, how could it be made permanent?

EXPERIMENT 79.—**Silver oxide.**—Add some sodium hydroxide solution to a dilute solution of silver nitrate. The precipitate is silver oxide, Ag_2O. Complete the equation—

$$2AgNO_3 + 2NaOH = Ag_2O + \ldots + \ldots$$

What experiment have you made upon the solubility of

gold? Is it soluble in nitric acid? In hydrochloric acid? In a mixture of both? How could you separate gold from silver in an alloy of both metals?

Questions

1. How could you convert silver nitrate into silver chloride? Silver chloride into silver nitrate?

2. How could you prepare pure silver from a silver coin containing 90 per cent silver and 10 per cent copper?

Problems

84. Calculate the percentage composition (*a*) of silver chloride, $AgCl$; (*b*) of silver sulphide, Ag_2S.

85. Calculate the formula of a compound of the following composition:

Silver	65.45
Sulphur	19.39
Arsenic	15.16

86. When hydrogen is heated with silver chloride silver is produced:

$$AgCl + H = HCl + Hg.$$

If 52.65 c.c. of hydrogen produce 0.505 gram of silver, what is the atomic weight of silver?

87. Calculate the percentage composition of silver acetate, $AgC_2H_3O_2$.

88. How much zinc is required to precipitate 5 grams of silver from solution?

$$2\overset{+}{Ag} + Zn = \overset{++}{Zn} + 2Ag.$$

CHAPTER XXI

No experiments.

Problems

89. How much auric chloride can be made from 65.67 grams of gold?

90. Calculate the formula of a compound containing 92.8 per cent of gold and 7.7 per cent of oxygen.

91. What volume of oxygen is produced when 49 grams of auric oxide are heated?

CHAPTER XXII

THE CALCIUM GROUP

EXPERIMENT 80.—**Calcium compounds.**—Support a piece of marble on a pipe-stem triangle and heat it with the flame of a blast-lamp for fifteen minutes. Examine the product. Is the change complete or partial? What gas has escaped? In what respects would the result have been different if you had heated the marble in a sealed vessel? Write the equation.

Place the mass in a beaker containing distilled water, allow it to stand for a time, and remove the unchanged marble. Decant the liquid carefully upon a filter. What is the name and formula of the white solid which is left? The liquid which runs through the filter is *lime-water*. Test its action upon red and blue litmus paper. Dissolve a little *ammonium oxalate* in distilled water and add it to some of the lime-water. The precipitate is calcium oxalate. This is a delicate test for calcium. Test faucet water carefully for calcium in the same way.

EXPERIMENT 81.—Examine calcium chloride. Expose a fragment to the air on a dry glass plate for an hour. (?) What is it used for in the laboratory? Make the flame test with it. Make the flame test with strontium chloride and barium chloride. (If the chlorates are at hand, use them instead of the chlorides. The result is more satisfactory.)

Make a very dilute solution of barium chloride, and add to it a few drops of sulphuric acid. The precipitate is *barium sulphate*, $BaSO_4$. This is a delicate test for $\overset{++}{Ba}$ ions, and conversely, a solution of barium chloride is a

delicate test for $\overset{++}{S}O_4$ ions; that is, for sulphuric acid or sulphates.

Divide the liquid containing the precipitate into two parts, and show that the latter is insoluble in hydrochloric acid and in ammonia water. This serves to distinguish barium sulphate from other precipitates of similar appearance.

Problems

92. A piece of pure marble weighing 10 grams is heated to complete decomposition. (*a*) What is the formula and weight of the substance which remains? (*b*) What gas escapes, and what volume measured at 20° and 740 mm.?

93. How many tons of limestone must be heated to make 200 tons of lime?

94. 1.363 kilos of marble are heated until entirely decomposed, and water is thrown on the residue. What is it, and what does it weigh?

95. Calculate the percentage composition (*a*) of barium sulphate, $BaSO_4$; (*b*) of barium carbonate, $BaCO_3$.

96. 10 grams of barium carbonate are dissolved in hydrochloric acid. (*a*) What volume of carbon dioxide is produced, and (*b*) how much crystallized barium chloride ($BaCl_2 2H_2O$) can be obtained from the solution?

97. How much barium sulphate can be made from 2 grams of calcium sulphate ($CaSO_4$)?

98. 1.182 grams of barium carbonate were dissolved in hydrochloric acid, and the solution precipitated with sulphuric acid. The barium sulphate obtained weighed 1.398 grams. Calculate the percentage of barium in the barium carbonate.

CHAPTER XXIII

MAGNESIUM

Experiment 82.—Examine *magnesium.* Is it light or heavy? Is its luster affected by the air? Hold a piece of magnesium ribbon 20 cm. (8 in.) long in forceps,

and burn it. Receive the product in a dish and reserve it. Burn a similar piece in steam, proceeding exactly as directed on page 173, Part I. Before introducing the magnesium, show that all the air has been expelled from the beaker by introducing a burning match or candle, which should be extinguished. Compare the products of burning magnesium in air and in steam. Are they identical? Investigate the action of hydrochloric, nitric, and sulphuric acids separately in magnesium. Cover a small piece of the metal with water in a test-tube, and add the acid gradually. Record the results. Remembering that magnesium is bivalent, write the equations for the action of hydrochloric and sulphuric acids upon it. The action of nitric acid is more complicated.

EXPERIMENT 83.—Carefully mix *on paper* 1 gram of powdered magnesium with 1.5 grams of powdered potassium chlorate. The materials must not be ground together. Place the mixture on a brick or a block of wood under the hood, stick a piece of filter paper about 10 cm. long in the heap, light the end of the paper farthest from the powder, and step aside. The action is explosive and the light intense. *If the burning paper fails to ignite the mixture, wait till you are absolutely sure that it is extinguished before approaching.*[1]

What are the products of the change? Why is the action so rapid? What application is made of mixtures of magnesium with substances yielding oxygen?

QUESTIONS

1. What is the valence of magnesium in $MgCl_2$, MgO, $MgSO_4$, and $Mg(NO_3)_2$? Why?

[1] It is safer and better to use paper which has been soaked in a solution of potassium nitrate and then dried. This burns without flame. It always ignites the mixture of magnesium and potassium chlorate.

2. Can you perceive any connection between the properties of magnesium oxide (almost infusible and non-volatile) and the great brightness of the flame of burning magnesium?

Problems

99. What volume of oxygen is needed to burn 9 grams of magnesium?

100. If 0.4 gram of magnesium liberated 391 c.c. of dry hydrogen at 18° when treated with HCl, what is the atomic weight of magnesium?

101. The electric current is passed through fused magnesium chloride until 14 grams of magnesium are obtained. What volume of chlorine at standard conditions is liberated?

CHAPTER XXIV

ZINC AND CADMIUM

Experiment 84.—Examine sheet *zinc* and record its properties. Refer to your notes of former experiments for data regarding the action of acids upon it. Heat a small piece of zinc upon charcoal with the blowpipe flame. Practise with the blowpipe until you can produce a continuous flame. The blast comes from the cheeks, not from the lungs, and the cheeks are refilled with air at intervals without interrupting the flame. What is the deposit upon the charcoal? Repeat Experiment 83, using zinc dust with the $KClO_3$ instead of magnesium. *Care.*

Experiment 85.—**Quantitative experiment. Atomic weight of zinc.**—Fit up the apparatus shown in Fig. 24. *T* is a graduated tube holding 100 c.c. It is filled with water, inverted in a vessel of water, and clamped. The flask, *F*, contains a piece of pure granulated zinc, which has been accurately weighed. It must not weigh more than 0.2 gram. A small piece of platinum wire is wound

around it to hasten the solution of the zinc. Before placing *D* under *T*, the doubly bored rubber cork should

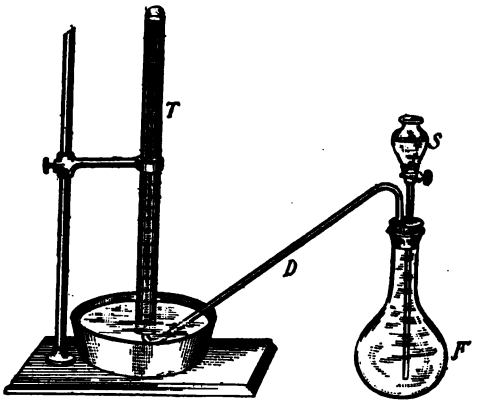

FIG. 24.

be twisted tightly into the flask, the end of *D* placed under water, and *F* and *D* filled with water by pouring water into the separating funnel *S*.

The stop-cock is then closed and the funnel filled with warm (*not hot*) dilute sulphuric acid, one part to three of water. *Remember the precautions necessary in diluting sulphuric acid.* The acid is cautiously allowed to run into *F*, and more of it admitted at intervals until the zinc is *completely dissolved.* The funnel must never become empty, or gas will escape. Remove *D*, cover the open end of *T* with the thumb, and transfer it to a cylinder of water which has had time to acquire the temperature of the laboratory. Clamp the tube so that the level of water inside and out is the same, place a thermometer in the water, and let the apparatus stand 15 minutes. Read the volume of the gas, the temperature, and the atmospheric pressure on the barometer in the laboratory. Record the readings, and repeat them. Take up the different steps of the calculation in the following order:

1. Find the volume your gas would occupy at 0°, 760 mm., and completely *dry*. The method of making the calculation is explained in the Appendix, and is the same that you employed in calculating the weight of a liter of oxygen (p. 38).

2. Calculate the weight of this gas by multiplying the corrected volume in *liters* by 0.0896, the weight of a liter of hydrogen under standard conditions.

3. Calculate by proportion the weight of zinc which would be required to set free 2 grams of hydrogen. Since zinc is bivalent, this will be the atomic weight of zinc.

EXPERIMENT 86.—Examine *cadmium,* and record its properties. Has it the same luster as zinc? Is it denser or lighter? Heat a chip of it on charcoal. (?) Dissolve a little cadmium chloride in water, and place a piece of zinc in the liquid. Allow to stand. What happens? Which of the two metals has the strongest tendency to exist as ions?

PROBLEMS

102. If 0.5 gram of zinc when dissolved in hydrochloric acid set free 183.7 c.c. of hydrogen measured over water at 15° and 760 mm., what is the atomic weight of zinc?

103. If 1 gram of zinc set free 366 c.c. of hydrogen measured over water at 9° and 748 mm., what is the atomic weight of zinc?

CHAPTER XXV

MERCURY

Mercury should not be brought into contact with jewelry. Rings should be removed before working with it.

EXPERIMENT 87.—Examine mercury. Is its luster affected by the air? Notice the high density (13.6). Place an iron nail and a fragment of marble upon the

surface of mercury in a beaker. Is there any other element which is a liquid at ordinary temperatures?

Heat a drop of it in a dry, clean test-tube, and describe the result. Incline the tube when introducing the metal, or the shock may break the tube.

Place a drop of mercury in a watch-glass and rub it upon a piece of clean copper. (?) What is an amalgam? Heat the copper carefully. (?)

Try the action of dilute hydrochloric, sulphuric, and nitric acids upon small drops of the metal.

Mercury must not be thrown into the sinks.

EXPERIMENT 88.—Place some mercuric sulphide, HgS, in a short tube about 0.5 cm. wide, open at both ends, and gently heat the portion containing the substance, inclining the tube so that a current of hot air shall be drawn over the mercuric sulphide. Describe and explain the result. Write the equation. Is this the usual effect of heating with abundant air-supply upon sulphides? What other sulphides might be expected to behave like mercuric sulphide under the same circumstances?

EXPERIMENT 89.—**Mercurous compounds.**—Add hydrochloric acid, drop by drop, to a solution of mercurous nitrate. The visible product is mercurous chloride. Does it look like silver chloride? Treat it with ammonia water. Does it behave like silver chloride with ammonia?

EXPERIMENT 90. — **Mercuric compounds.** — *Mercuric chloride is intensely poisonous. It must not be touched with the fingers.* Dissolve a little mercuric chloride in about 50 c.c. of water. To a portion of the solution add a solution of sodium hydroxide. The precipitate is mercuric oxide, HgO—yellow because it is finely divided. Complete the equation—

$$HgCl_2 + 2NaOH =$$

To another portion of the solution add a solution of potassium iodide, one drop at a time, shaking constantly.

The precipitate is mercuric iodide. Notice that the yellow modification forms first, and immediately passes into the red stable modification. Write the equation. Show that the precipitate is soluble in more potassium iodide solution.

To the rest of the solution add an equal volume of a solution of stannous chloride, $SnCl_2$, and heat the liquid. Explain.

Question

Make a tabular statement of the differences between mercurous and mercuric chlorides, using your experiments and the text of Part I as sources of information.

Problems

104. 758 grams of mercuric chloride are dissolved in water. (*a*) How much potassium iodide must be added to the liquid, and (*b*) how much mercuric iodide will be obtained?

105. If mercurous chloride contains 84.93 per cent mercury and 15.07 per cent chlorine, and if the formula is Hg_2Cl_2, what is the atomic weight of mercury?

106. If mercuric chloride has the formula $HgCl_2$ and contains 73.8 per cent of mercury and 26.2 per cent of chlorine, what is the atomic weight of mercury?

107. 88.5832 grams of mercuric sulphide, when completely decomposed, yield 76.3725 grams of mercury. What is the atomic weight of mercury?

CHAPTER XXVI

BORON AND ALUMINIUM

Experiment 91.—**Boric acid and borax.**—Dissolve about 10 grams of powdered borax in 50 c.c. of hot water in a beaker. Add 25–30 drops of strong sulphuric acid, and allow to cool. Boric acid, $B(OH)_3$, crystallizes.

Examine boric acid from the stock bottle. Hold a

fragment of it in the Bunsen flame. The flame color can be used as a test for boric acid. Place a little borax (0.2 gram or less) in a dish and moisten it with strong sulphuric acid. Stir with a glass rod, add 5 c.c. of alcohol, and stir again. Set fire to the mixture. Record the result.

Make a little loop in the end of a piece of platinum wire, heat the loop red hot, and dip it into a little borax. Put the wire with the borax which adheres to it back into the flame. What is the cause of the swelling up of the borax? What is the name and composition of the transparent bead which remains?

Heat the bead, and cause a minute quantity of *manganese dioxide* to adhere to it. Melt the bead again, and let it cool. The color is characteristic of *manganese.* Slowly turn the cap at the base of the burner until a small portion of the flame is faintly luminous, and hold the bead in this portion steadily. When cold it should be colorless. Heat it in the outer portion of the flame to restore the color.

Remove the bead by dipping it while hot into cold water and scraping it off, and repeat this experiment, using the blowpipe instead of the burner flame. Use a blowpipe-tip on your burner, closing the holes at the base. The *oxidizing flame* of the blowpipe should be used for making the borax bead, containing the manganese. This flame corresponds to the outer non-luminous portion of the Bunsen-burner flame.

The oxidizing flame is produced by placing the end of the blowpipe nearly in the middle of the flame and blowing steadily. The flame should be blue, *perfectly noiseless,* and parallel to the slit in the burner-tip.

The *reducing flame* should be used to decolorize the manganese borax bead. To produce it, place the tip of the blowpipe at the edge of the flame, and use a more

gentle current of air than that used for the oxidizing flame. The flame will be partly blue and partly yellow. It must be noiseless and steady. It should surround the bead completely, so as to prevent the access of air.

Remove the bead. Make a fresh one, and investigate the color produced in it by a minute fragment of *copper sulphate* in both the outer and inner flames. Record the result.

Make another bead, take up with it a scarcely visible speck of *cobalt nitrate,* and study the color produced.

What use can be made of these phenomena in testing for the metals?

EXPERIMENT 92.—**Aluminium.**—Examine sheet aluminium. Is it affected by the air? How does its density compare with that of familiar metals, like iron and copper? How do small pieces of it behave with *dilute* sulphuric, nitric, and hydrochloric acids? Try each acid separately. What gas is given off? What is the action of a solution of sodium hydroxide on a piece of aluminium?

Examine aluminium powder. This is the "bronze paint" used for mail boxes. Place a heap of it 1 cm. wide on an asbestos plate, and set fire to it. What effect has aluminium upon metallic oxides at high temperatures, and what use is made of the fact?

EXPERIMENT 93.—**Alum.**—Dissolve 10 grams of *aluminium sulphate* in a *little* hot water. In another beaker dissolve 5 grams of *potassium sulphate* in a small quantity of hot water. Both liquids must be clear. If not, decant or filter. Mix the liquids in a glass dish or large beaker, and allow to cool. The product is *potassium-alum.* Place some of the crystals on a glass plate, and examine them with a lens.

QUESTIONS

1. What are some of the uses of aluminium?
2. Why are there so many different alums? What chemical

composition must a substance have in order to be called an alum?

3. Why is boron considered a non-metal and aluminium a metal?

Problems

108. Calculate the percentage composition of borax, $Na_2B_4O_7 \cdot 10H_2O$. Calculate *water*, not hydrogen.

109. What is the formula of a substance containing 31.19 per cent boron and 68.81 per cent oxygen?

110. If 6.75 grams of aluminium, when dissolved in hydrochloric acid, yield 8.4 liters of hydrogen, what is the atomic weight of the metal?

CHAPTER XXVII

SILICON

Experiment 94.—**Hydrogen silicide.**—Double a piece of magnesium ribbon 6 cm. long, and heat it in a sealed tube of hard glass to a bright-red heat for several minutes. Allow to cool, break open the tube, discard the magnesium if any remains, and throw the blackened fragments of glass into dilute hydrochloric acid, 1 part to 2 parts of water, in a beaker. The magnesium produces *magnesium silicide,* $SiMg_2$, with the SiO_2 of the glass. This liberates hydrogen silicide with the acid. Write all the reactions, and explain the cause of the spontaneous ignition of the gas.

Experiment 95.—Add strong hydrochloric acid to a solution of sodium silicate in a porcelain dish. What is the composition of the jelly which separates? Evaporate slowly to dryness under the hood, allow to cool, moisten the residue with hydrochloric acid, half fill the dish with water, heat to boiling, and filter. The substance left on the filter is *silicon oxide,* SiO_2. Wash it with water until a few drops of the liquid which runs through gives no

precipitate when collected in a test-tube, and mixed with a drop of silver nitrate solution.

Examine some crystals of quartz. Make a drawing of one in your note-book. Can you scratch quartz with a knife? With a file? Is a quartz crystal hard enough to scratch glass?

Questions

1. In what respect are silicic acid and nitrous acid similar?

2. What is meant by the term *silicate*?

Problems

111. What is the weight of 2.8 liters of hydrogen silicide under standard conditions?

112. Calculate the formula of a silicate which was found by analysis to possess the following composition:

Zinc	58.6 per cent.
Silicon	12 7 "
Oxygen	28.7 "

CHAPTER XXVIII

TIN

Experiment 96.—Examine tin-foil, bar tin, and granulated tin, and record the properties of the metal. Are ordinary "tin" vessels made of solid tin? Investigate this by bringing a magnet in contact with a tin vessel of any kind. Tin is not attracted by the magnet. Try this with tin-foil.

Study the effect of heat upon tin by placing some tin-foil in a hollow scraped in a piece of charcoal and letting the blowpipe flame play upon it. Does it melt? Does it absorb oxygen? Place a drop of mercury in a watch-glass and press some tin-foil beneath the surface. What is the product?

EXPERIMENT 97.—Heat some tin-foil under the hood for some time with strong hydrochloric acid in a wide test-tube. The result is a solution of stannous chloride, $SnCl_2$. Dilute the liquid, and place a piece of zinc in it. The product is *tin*. Describe it, and notice that a finely divided metal may have a very different appearance from the compact substance. Have you noticed any other examples of this fact?

EXPERIMENT 98.—**The atomic weight of tin.**—Clean carefully a porcelain crucible, dry and weigh it. Place a little granulated tin in it and weigh again to ascertain the amount of tin taken. This should be from 0.5 to 0.7 gram. The quantity must be accurately known. Cover the bottom of the crucible with water and slowly add nitric acid until the tin is covered. Support the crucible on a pipe-stem triangle, and place a small flame 5 cm. below the bottom. This should be done under the hood. When the action seems complete, evaporate cautiously to dryness, avoiding any spattering which will cause loss of substance and make it necessary to go back to the beginning. When the residue is dry, heat it, gradually at first, and finally intensely; let it cool, and weigh it. There will be a gain in weight, which is due to the fact that the tin now exists as SnO_2. Calculate the atomic weight of tin thus—

Gain in weight : O_2 = weight of tin : x.

For example:

Weight of crucible + tin.....	16.642
Weight of crucible, empty	14.642
Weight of tin	2.000
Weight of crucible + SnO_2.....	17.180
Weight of crucible + tin.....	16.642
Increase due to oxygen.....	.538

$$0.538 : 32 = 2 : x.$$

$$x = 119.$$

QUESTIONS

1. Discuss the properties of gray tin and its relations to ordinary tin.

2. If 20° is the temperature at which gray tin and white tin are in equilibrium, what is the condition of the tin of a roof on a winter day? How is it that tin can be used for any purpose requiring exposure to cold?

PROBLEM

113. 1 gram of tin was treated with nitric acid. After evaporation to dryness and heating, the gain in weight was 0.271 gram. Calculate the atomic weight of tin.

CHAPTER XXIX

LEAD

EXPERIMENT 99.—Examine *lead.* Cut a chip off a piece in order to observe the luster of the untarnished metal. Heat a little granulated lead in a test-tube with dilute nitric acid. The product is *lead nitrate.* Examine *litharge,* PbO, and heat a little of it on charcoal with the blowpipe flame. Describe and explain the result. Examine *lead dioxide,* PbO_2. Heat a little of it in a tube sealed at one end (hard glass). Apply the spark test to the gas given off. What is the residue? Examine *red lead.* Treat it with dilute nitric acid. What is the cause of the color-change?

EXPERIMENT 100.—Dissolve about 5 grams of lead nitrate in 100 c.c. of water. Use small portions of this liquid to study the behavior of lead salts with the following substances:

	Product.
Sulphuric acid, H_2SO_4	Lead sulphate, $PbSO_4$
Potassium chromate, K_2CrO_4	Lead chromate (chrome yellow), $PbCrO_4$
Hydrogen sulphide, H_2S	Lead sulphide, PbS

Write all the equations. In the remainder of the lead nitrate solution place a piece of zinc, and allow it to remain undisturbed as long as possible. The product is *lead.*

Questions

1. What is the effect of the continual absorption of small quantities of lead upon the body, and what are the antidotes?

2. What is the valence of lead in PbO, $PbCl_2$, PbS, $Pb(NO_3)_2$, and PbO_2?

Problems

114. If 50 grams of litharge (PbO) contain 3.5862 grams of oxygen, what is the atomic weight of lead?

115. What are the name and formula of a compound of the following composition?

Lead	77.52 per cent.
Carbon	4.49 "
Oxygen	17.98 "

CHAPTERS XXX AND XXXI

No experiments.

Questions

1. What reasons have we for the belief that red and colorless phosphorus are two forms of the same element?

2. Discuss the luminosity of colorless phosphorus. How is the luminosity affected by the composition of the gas in contact with the substance?

Problems

116. How many liters of oxygen are needed to burn 93 grams of phosphorus to P_2O_5?

117. 20 grams of phosphorus are burned in a vessel from which nothing is allowed to escape. How much will the vessel increase in weight?

118. What volume of air is needed to burn 124 grams of

phosphorus to P_2O_5? Assume that air contains 21 per cent by volume of oxygen.

119. If 4 grams of phosphorus when burned yield 9.16 grams of P_2O_5, what is the atomic weight of phosphorus?

CHAPTERS XXXII AND XXXIII

ARSENIC AND ANTIMONY

EXPERIMENT 101.—**Arsenic.**—Examine the element and record its properties. Heat a fragment of it in a tube of hard glass, closed at one end. Does it melt? Does it vaporize? The two sublimates (steel-gray and black) are different allotropic forms of arsenic. Heat a fragment of arsenic the size of the head of a pin on charcoal with the blowpipe flame. The product is *arsenious oxide.*

EXPERIMENT 102.—*Arsenious oxide.*—Examine the substance and compare it with the *element* arsenic. Heat a trace of arsenious oxide in a sealed tube. (?) Examine the sublimate with a lens. What is the shape of the crystals?

Dissolve about 0.1 gram of arsenious oxide by boiling it gently in a test-tube with dilute hydrochloric acid. Dilute the liquid to 100 c.c. Into a portion of this liquid pass hydrogen sulphide gas—*a slow succession of bubbles.* Use a Kipp generator as a source of the gas, or generate it in a wide test-tube, as directed in Experiment 110. *Work under the hood, and do not inhale the hydrogen sulphide.* Describe what takes place, and write the equation.

Place a clean piece of sheet copper about 2 cm. square in a dish, fill the dish with water, add about 1 c.c. of hydrochloric acid, and heat almost to boiling. Is there any action? Add 1 c.c. of your arsenic solution to the

liquid, and continue heating for ten minutes. The deposit on the copper is *arsenic*. Remove the copper, dry it carefully with filter paper, roll it up, and place it in a tube of hard glass, sealed at one end. Heat it gently. (?) Look for a sublimate with a lens.

Mix about 0.1 gram of arsenious oxide with twice its weight of powdered charcoal, and heat the mixture in a tube of hard glass, sealed at one end. Introduce the mixture with a paper trough. The upper part of the tube must be *clean*. Describe and explain the result.

EXPERIMENT 103.—**Antimony.**—Examine the element. Is it a metal? How does it behave under the hammer? Heat a fragment on charcoal with the blowpipe. The product is *antimonious oxide*, Sb_4O_6. Let the globule of melted antimony fall while still red hot on a sheet of paper. (?)

QUESTION

Explain in detail what is meant by the statement that arsenic and antimony stand on the border-line between metals and non-metals. Consider both physical and chemical properties in your answer.

PROBLEMS

120. How much charcoal must be used to reduce 132 grams of arsenious oxide? What volume of carbon monoxide will be given off?

$$As_4O_6 + 6C = 6CO + As_4.$$

121. In a case of poisoning the arsenic from the body was converted into arsenious sulphide, As_2S_3, and weighed. The weight of the As_2S_3 was 0.82 gram. What quantity of arsenious oxide had been administered?

122. What is the weight of 5.6 liters of arsine at 0° and 760 mm.?

123. Calculate the formula of a compound of the following composition:

Arsenic	48.39 per cent.
Sulphur	51.61 "

CHAPTER XXXIV

CHROMIUM

EXPERIMENT 104.—**Chrome alum. Preparation of potassium chromate, K_2CrO_4.**—Examine chrome alum. What is the form of the crystals? Powder a little. With a portion of the powder try the solubility in water. What is the color of the liquid? Chrome alum contains *chromium sulphate,* $Cr_2(SO_4)_3$, and in it and the other chromium *salts* chromium plays the chemical *rôle* of a trivalent metal, like aluminium.

Mix the rest of the powdered chrome alum with potassium nitrate and potassium carbonate in a clean mortar. The mixture should contain equal quantities of the three substances. Fuse the mixture on a piece of platinum foil, held in forceps. Notice the color of the melted substance. It now contains *potassium chromate,* K_2CrO_4, in which the chromium plays the *rôle* of a non-metal, like sulphur.

Dissolve the melted substance in water and carefully add acetic acid until litmus paper is turned red. Notice the color of the liquid. It is due to the negative ion $\overset{-}{C}r\overset{-}{O}_4$. What color did positive ions $\overset{+++}{Cr}$ give to water when you dissolved chrome alum in it? Divide the liquid into two portions. To half of it add solution of lead nitrate. Complete the equation—

$$K_2CrO_4 + Pb(NO_3)_2 = PbCrO_4 + \ldots$$

Lead chromate

To the rest of the liquid add either nitric or sulphuric acid, drop by drop, until the yellow changes to red. The red liquid contains *potassium dichromate,* $K_2Cr_2O_7$. Complete the equation—

$$2K_2CrO_4 + 2HNO_3 = K_2Cr_2O_7 + \ldots + \ldots$$

EXPERIMENT 105.—**Preparation of potassium chromate from potassium dichromate.**—Dissolve 5 grams of potassium dichromate in water in a porcelain dish. Add a solution of potassium hydroxide until the liquid is yellow. Evaporate to small volume and let the solution crystallize. The product is *potassium chromate.* Complete the equation—

$$K_2Cr_2O_7 + 2KOH = \ldots + \ldots$$

EXPERIMENT 106.—**Preparation of chromic chloride from potassium chromate.**—Powder a *little* potassium chromate, and heat it gently in a test-tube under the hood with strong hydrochloric acid. What gas escapes? What is the change in the color of the liquid? It now contains *chromic chloride,* $CrCl_3$, in which chromium acts like a metal, chemically. Supply the lacking numbers in the equation—

$$K_2CrO_4 + 8HCl = CrCl_3 + KCl + Cl + H_2O$$

QUESTIONS

1. Briefly state the steps by which you have transformed a chromic salt into a chromate. A chromate into a chromic salt.

2. How can potassium chromate be converted into potassium dichromate? Potassium dichromate into potassium chromate?

PROBLEM

124. Calculate the percentage composition of *chromite,* Cr_2O_3FeO.

CHAPTER XXXV

SULPHUR

EXPERIMENT 107.—**Different forms of sulphur.**—Examine *roll sulphur.* Is it brittle or malleable? dense or light? Rub a piece on the coat-sleeve and bring it near

small pieces of paper. Does it conduct electricity? Hold a piece firmly in the hand close to the ear and notice the crackling noise caused by the portions of sulphur next the hand, expanding as they become warm and cracking away from the other portions, which remain cold. This shows that it is a bad conductor of heat.

Examine *flowers of sulphur,* and record its properties. Powder about 2 grams of roll sulphur and shake it up with 3–5 c.c. carbon disulphide in a dry test-tube. *Caution: Keep the carbon disulphide bottle corked, and avoid the vicinity of flame.* When the sulphur has dissolved, pour the liquid into a dish and let it evaporate spontaneously. Examine the crystals with a lens. They consist of *α-sulphur.*

Fill the smallest size Hessian crucible with crushed roll sulphur, place it in a ring of your stand, and apply heat until the sulphur has melted. Allow to cool. Just when the surface begins to solidify take the crucible in forceps and pour out the liquid portion into a pan of water. Let it cool, and examine the interior. It is covered with crystals of β-sulphur. Make a drawing of a mass of the crystals. Record their properties. Notice that they are different in color as well as shape from the crystals of α-sulphur. Preserve some crystals overnight and explain the change which takes place.

When roll sulphur is first made, does it consist of α-sulphur or β-sulphur? Which does it consist of after being preserved for a time?

In what respects does the change of α-sulphur into β-sulphur by heat (p. 236, Part I) resemble, and in what respects differ, from the change of ice into water?

EXPERIMENT 108.—**Fusion of sulphur. Soft sulphur.** —Fill a wide test-tube about one-quarter with crushed roll sulphur, hold it with a paper holder, and heat slowly. Obtain the three states of fusion described on page 235,

Part I. The three conditions should be perfectly distinct. Let the sulphur in the second liquid stage cool slowly, and note that it passes through the changes in the reverse order.

Now heat the sulphur until it boils vigorously, and pour the liquid in a thin stream into a bottle of cold water, lighting the sulphur vapor at the mouth of the tube while pouring. Examine the product and describe it. Preserve some of it. (?) Is it a stable condition of sulphur?

EXPERIMENT 109.—**Union of sulphur with metals.**—Fill the same test-tube one-quarter with sulphur, and heat to boiling. Into the tube throw a little powdered iron. (?) Straighten a thin iron wire, heat one end of it to redness, and quickly place it in the sulphur vapor. Cut a thin strip of copper and hold it in the tube.

How does sulphur behave with metals? What is a *sulphide?*

EXPERIMENT 110.—**Hydrogen sulphide.**—*Use the hood. Hydrogen sulphide is poisonous, and must not be inhaled.* Cover the bottom of a gas-generating bottle like the one used in making hydrogen with lumps of *iron sulphide,* FeS. Incline the bottle, and slide the solid in slowly. Cover with water, and add hydrochloric acid through the funnel-tube. Collect several wide test-tubes full of the gas over water as warm as you can work with. (Why is it necessary to use *warm* water?) Cork one of the test-tubes, transfer it to a pan of cold water, remove the cork, and shake the tube, keeping its mouth under water. Is the gas soluble in water? With another test-tube try its combustibility. Meanwhile, allow the hydrogen sulphide to bubble through a bottle or beaker containing at least ½ liter of cold water. Hold the exit tube of your apparatus near a silver coin. A clean piece of sheet copper. A paper wet with a solution of lead acetate. Record and explain the results.

Put the exit tube back in the water and continue passing the gas through for ten minutes. Meanwhile, prepare very dilute solutions of copper sulphate, lead nitrate, cadmium chloride, tartar emetic (an antimony compound), and sodium chloride. Stop the production of the gas by filling the bottle with water and pouring away the liquid. Has the water in the beaker the odor of the gas? Use it to study the effect of hydrogen sulphide upon the solutions you have prepared. Complete the equations:

$$CdCl_2 + H_2S = CdS +$$
$$Pb(NO_3)_2 + H_2S = PbS +$$
$$CuSO_4 + H_2S = CuS +$$

Does the hydrogen sulphide affect the sodium chloride solution? Other metals—e. g., potassium, calcium, and magnesium—behave like sodium. How could you separate one of these metals from lead, cadmium, or copper?

Questions

1. What is the natural state of sulphur at ordinary temperatures? At 100°? Give facts to support your answer.

2. Why is hydrogen sulphide so much used in analytical chemistry?

Problems

125. 2 grams of crystallized copper sulphate ($CuSO_4 5H_2O$) are dissolved in water, and it is required to precipitate all the copper as CuS. How much iron sulphide, and how much hydrochloric acid containing 25 per cent HCl, are needed to generate enough H_2S for the purpose?

126. (*a*) What volume of hydrogen sulphide is produced when 17.6 grams of FeS are dissolved in HCl? (*b*) From what weight of mercuric chloride dissolved in water will this quantity of H_2S precipitate the mercury as mercuric sulphide?

$$HgCl_2 + H_2S = HgS + 2HCl.$$

127. What volume of air is needed to burn 500 grams of sulphur to SO_2? Assume that air contains 21 per cent of oxygen by volume.

128. What volume of H_2S escapes when 5 grams of iron sulphide are dissolved in HCl?

129. How much iron sulphide is needed to make 50 liters of H_2S?

In working with sulphur dioxide, use the hood. The gas is irritating, and must not be inhaled.

EXPERIMENT 111.—**Sulphur dioxide, SO_2.**—Fill a small flask one-third with a strong solution of *sodium acid sulphite,* $NaHSO_3$, and allow strong sulphuric acid to drop slowly into it from a separating funnel which passes through one hole in the cork. Lead away the SO_2 by a tube passing through the other hole, and collect it by downward displacement in dry, covered vessels. Collect two wide test-tubes and two bottles of the gas. Cover the test-tubes with paper and the bottles with glass plates.

Investigate its solubility by placing a test-tube filled with it mouth down in water. Remove the test-tube from the pan by slipping a glass plate under it, and add a few drops of litmus solution to the liquid in the tube. The reddening of the litmus shows the presence of *sulphurous acid.* How has it been formed?

Place a few drops of litmus solution in a test-tube filled with the gas, cork the tube, shake it, and let stand. What is the first action of the gas on the litmus? Why? What is the final action? Sulphur dioxide is much used for bleaching.

Lower a lighted candle into a bottle of the gas. (?) Place some *lead dioxide* in a clean deflagrating spoon and warm it gently. Lower it into a bottle of sulphur dioxide. The product is *lead sulphate*—

$$PbO_2 + SO_2 = PbSO_4$$

Examine it, and compare it with lead dioxide.

Pass sulphur dioxide from your generator through 100 c.c. of water, until you have a strong solution. Test the liquid with litmus paper (red and blue). Add a

dilute solution of sodium hydroxide to it drop by drop, stirring constantly until litmus paper is just turned blue. Then evaporate to dryness in a clean dish. The residue is *sodium sulphite,* Na_2SO_3, formed thus—

$$2NaOH + H_2SO_3 = Na_2SO_3 + 2H_2O$$

Examine it. How does it behave with hydrochloric acid? with sulphuric acid? What gas is produced when the acids act upon it? Take some sodium sulphite from the bottle and see if it behaves in the same way with acids.

Most sulphites are insoluble in water. Mix water solutions of sodium sulphite and calcium chloride. The visible product is calcium sulphite, $CaSO_3$. Write the equation. Is it soluble in water? in hydrochloric acid?

If time permits, pass a slow current of sulphur dioxide from your generator through a U-shaped tube, which is embedded in a mixture of equal parts of ice and salt. Liquid sulphur dioxide collects. Pour some of it on a *little* water in a porcelain dish. *Use the hood.* The water will be frozen by the rapid evaporation of the sulphur dioxide.

EXPERIMENT 112.—**Sulphuric acid.**—Examine sulphuric acid. How does it compare with water as regards density and consistency? Slowly pour sulphuric acid in a thin stream, stirring constantly into 5 c.c. of water in a beaker until you have added an equal volume of the acid. *Care. Never add water to sulphuric acid. Always add the acid slowly to the water, stirring constantly.* Notice the production of heat. Make some dilute sulphuric acid, and write a word with it on a sheet of paper, using a glass rod as a pen. Heat the paper over the flame, taking care not to set fire to it. Place a splinter or a match-stick in a little strong sulphuric acid in a test-tube.

Make a *little* strong hot solution of sugar and place not more than 3 c.c. of it in a test-tube. Slowly add strong sulphuric acid until a decided result is obtained.

Wood and paper are chiefly composed of a substance called cellulose, whose formula is $C_6H_{10}O_5$. The formula of sugar is $C_{12}H_{22}O_{11}$. Bearing in mind the strong attraction of sulphuric acid for water, explain the three results just obtained.

Repeat the reaction of barium chloride with sulphuric acid. (Experiment 81.)

Look through your notes from the beginning for all cases in which sulphuric acid has been used, and make a table in your notes containing the equations for the various reactions in which you have employed it. Arrange the table thus:

Reactions of Sulphuric Acid with Various Substances

Zinc.....	$Zn + H_2SO_4 = ZnSO_4 + H_2$
Sodium nitrate.....	$2NaNO_3 + H_2SO_4 = Na_2SO_4 + 2HNO_3$

Questions

1. What is the most recent process for the production of sulphuric acid?

2. What is the lead-chamber process for the production of sulphuric acid?

Problems

130. How much sulphuric acid can be made from 4 tons of sulphur?

131. How much sulphuric acid can be made from 40 tons of pyrite, FeS_2?

132. The density of sulphuric acid is 1.84. How much sulphur is there in 100 c.c. of it?

133. How many tons of sulphuric acid can be made from 100 tons of pyrite containing 48 per cent of sulphur?

134. 1.8752 grams of cobalt, when converted into cobalt sulphate, yielded 4.9472 grams. What is the atomic weight of cobalt? Assume S = 32, O = 16. The formula of cobalt sulphate is $CoSO_4$.

CHAPTER XXXVI

No experiments.

CHAPTER XXXVII

FLUORINE

EXPERIMENT 113.—**Hydrofluoric acid.**—Mix in a lead dish—*under the hood*—some powdered calcium fluoride with about an equal weight of strong sulphuric acid. Support the dish in a ring or on a pipe-stem triangle, and apply a gentle heat. The flame should not touch the dish. *Do not inhale the gas given off. It is poisonous.* Hold over the dish a piece of red and a piece of blue litmus paper. (?) Wet a glass rod with ammonia. Be careful not to let any ammonia fall into the dish, since this would cause an explosion. Remembering the behavior of ammonia with HCl, explain its behavior with HF. Cover the dish with a clean plate of window-glass, and let it stand some hours, if possible. Remove the glass, wash it with water, and examine. Clean the dish by scraping the contents into a jar with an old knife or file.

Warm a glass plate by placing it on a ring of your stand and moving about a small flame at least 5 cm. below the plate. If the flame is brought nearer the plate it will crack it. Place some paraffin on the warm plate, and allow the melted paraffin to cover it completely. Let the plate cool and scratch a word or some lines through the wax with a pin or a knife-blade. Take the plate to

the hood, cover the word with a piece of filter paper, and carefully wet the paper with hydrofluoric acid. *Be careful. Hydrofluoric acid produces poisoned wounds on the skin, and its vapor is poisonous.* Place a paper containing a plainly written caution by the side of the plate, so that no one may touch it and be injured, and let it stand under the hood for as long a time as possible. Then take away the paper *with forceps,* and throw it into the waste jar. Take the plate with forceps and place it in a pan of boiling water. This will remove both the hydrofluoric acid and the wax. Wipe it with a towel and examine. Write the equations between calcium fluoride and sulphuric acid and between silica and hydrofluoric acid.

Problems

135. How much calcium fluoride and how much sulphuric acid containing 96 per cent H_2SO_4 are needed to make 12 grams of pure hydrofluoric acid?

136. How much calcium sulphate and how much hydrofluoric acid are formed when 50 grams of calcium fluoride are heated with sulphuric acid?

CHAPTER XXXVIII

BROMINE AND IODINE

Experiment 114.—**Bromine.**—*Use the hood in working with bromine. The vapor is extremely irritating. The liquid must not be gotten upon the flesh or clothing.* Powder 1 gram of potassium bromide and mix it with 2 grams of powdered manganese dioxide. Place the mixture in a wide test-tube, and fill the tube one-quarter with cold dilute sulphuric acid, 1 volume acid to 4 volumes water. Clamp the tube in an inclined position, connect it with a *well-fitting* cork and delivery tube, dipping into

an empty test-tube surrounded by cold water. Heat gently. The product is *bromine.* Examine it. The equation is similar to that given for *iodine,* on p. 258, Part I. Write it.

Examine bromine, attending especially to color, density, and volatility. Place three dry test-tubes in a rack *under the hood.* Put a little bromine in each (1 cm. deep). Notice the color of the vapor. Into one test-tube drop a piece of tin, into another a piece of antimony. Use pieces a little larger than the head of a pin, and be careful. Is bromine an energetic element. Nearly fill the third test-tube with water and cork it tightly. Shake it. Notice the color of the product, which is called *bromine-water.* Does bromine-water bleach? Try it on a little water faintly colored with litmus.

Dissolve a little potassium bromide in water in a wide test-tube. Add a layer of chloroform 2 cm. deep, cork the tube, and shake it. Do bromine *ions* give any color to water or chloroform? Now add *chlorine-water* to the liquid until the tube is nearly full, cork it, and shake it for five minutes. The color of the chloroform is due to dissolved bromine, Br_2, not bromine ions. Write the equation, first in the ordinary way; second, using the idea of ions. The chloroform and the water take no part in the chemical change, and must not appear in the equation.

Pour away the water as far as possible without losing any of the chloroform, fill up the tube with water, shake and pour off again to get rid of all chlorine. A solution of Br_2 in chloroform remains in the tube. Add more water and a crystal of potassium iodide, KI, and shake again. Notice the change in the color of the chloroform. It now contains, instead of Br_2, I_2 dissolved. Write the equation, both from the ordinary standpoint and from the standpoint of ions.

Discuss this experiment at length in your notes, and

show from it that the tendency of these three elements to exist as ions decreases in the order Cl, Br, I. Where would F come in this series?

EXPERIMENT 115.—**Hydrobromic acid.**—Place a few crystals of potassium bromide in a wide test-tube and cover with a mixture of 3 volumes of sulphuric acid and 1 volume of water. Clamp the tube in an inclined position, insert a cork with a delivery tube, and collect the gas by downward displacement in *dry* wide test-tubes. Use a gentle heat. Try the solubility in water. Try the action upon red and blue litmus paper. Let fall a drop of ammonia into a tube of the gas. Make a comparison between HBr and HCl.

EXPERIMENT 116.—**Iodine.**—Place in a mortar 1 gram of potassium iodide and 2 grams of manganese dioxide. Grind thoroughly and transfer the mixture to a porcelain dish. Add enough of a cold mixture of 1 volume strong sulphuric acid and 2 volumes water to cover the powder. Cover the dish with a funnel and apply a *gentle* heat. When the funnel is well coated with iodine-crystals, stop the experiment and scrape off the crystals. Use them for the following experiments, or, if the quantity is insufficient, obtain a little iodine from the bottle.

Examine iodine with respect to color, luster, odor, and density. Warm a dry wide test-tube *gently,* and drop a crystal of iodine into it. Invert the tube. What are the properties of iodine vapor? What about its density? If the formula of iodine is I_2, what must be the density of its vapor referred to hydrogen? referred to air?

Use the same test-tube to test the solubility of iodine in water. (?) Use the tip of a knife-blade filled with powdered iodine—not more. Add two crystals of potassium iodide and shake. (?) Preserve the solution.

Grind about 0.5 gram of starch with a little water in a mortar. Slowly pour the liquid into 100 c.c. of boiling

water in a clean dish. The liquid must be colorless. It remains colorless when a few drops of it are mixed with a little potassium-iodide solution in a test-tube. Try it. Now half fill a wide test-tube with water, add a little of your starch solution, and a drop of your solution of iodine. Notice the color, which can be used as a test both for iodine and starch. The color disappears when the tube is heated, but reappears on cooling. Try this.

Place about 3 c.c. of chloroform in a test-tube, half fill the test-tube with water, and add a drop or two of your solution of iodine in potassium iodide. Shake the tube. The color of the chloroform solution can be used as a test for iodine.

Carefully examine potassium iodide. What are the corresponding bromine and chlorine compounds? Does it resemble them? Add a few drops of potassium iodide solution to a little solution of silver nitrate. The visible product is *silver iodide,* AgI. Write the equation. Make some *silver bromide* in the same way. Do they resemble silver chloride? If you can not recall the properties of silver chloride, make a little by mixing solutions of silver nitrate and sodium chloride. Use half of each product to study the solubility of the three substances in ammonia. Stand the other half in direct sunlight, or the brightest light accessible. Record and discuss the results.

Problems

137. How much potassium bromide is required to make 50 grams of bromine?

138. 1.8 grams of potassium bromide are heated with dilute sulphuric acid. What volume of hydrobromic-acid gas at 26° is liberated?

139. What volume of hydrobromic-acid gas can be made from 160 c.c. of liquid bromine according to the method described on p. 257, Part I? Assume that bromine is 3 times as dense as water.

140. Manganese dioxide is heated with hydrochloric acid, and the chlorine passed in a solution of potassium iodide. How much iodine will be set free by the chlorine evolved when 12 grams of manganese dioxide are used?

141. Under the same circumstances as in Problem 138, how much manganese dioxide is needed to liberate 63.5 grams of iodine?

142. Under the same conditions as in the two preceding problems, how much iodine will be set free when 43.5 grams of manganese dioxide are used?

CHAPTER XXXIX

IRON

EXPERIMENT 117.—**Reduction of ferric oxide by hydrogen.**—Place a little ferric oxide in a bulb-tube and pass hydrogen, dried by passing through a U-tube filled with lumps of calcium chloride, over it. It is best to use a Kipp generator. If you employ an ordinary gas-bottle, it is necessary to wait until all air is expelled, and to wrap the gas-bottle and cork in a towel before applying heat to the bulb. Heat the bulb, gently at first. What are the two products? Let the bulb cool completely, shake out the solid material on a paper, and examine it. Is it attracted by the magnet? Is it combustible?

EXPERIMENT 118.—**Spontaneous oxidation of finely divided iron.**—Heat in a sealed tube of hard glass some *ferrous oxalate* to redness, keeping the finger loosely over the open end of the tube to prevent the entrance of air. Cork the tube tightly, let it cool, and shake out the black iron-powder into a plate or a dry dish. (?) Clearly, finely divided iron oxidizes more readily than compact iron. You have noticed frequently in your work that powdering a substance makes it dissolve more rap-

idly. The reason is the same in the two cases. What is it?

Examine the most important ores of iron, hematite (Fe_2O_3), magnetite (Fe_3O_4), and limonite, which contains both ferric oxide and ferric hydroxide. Also examine pyrite (FeS_2), which is important on account of its sulphur.

EXPERIMENT 119.—**Ferric salts.**—Dissolve a little *ferric chloride,* $FeCl_3$, in 50 c.c. of water in a beaker. Use small portions to investigate the action of the following substances upon it:

1. Ammonia water, NH_4OH, precipitates ferric hydroxide, $Fe(OH)_3$. Write the equation and record the properties of the substance.

2. *Potassium thiocyanate,* KCNS (often called potassium sulphocyanide). Dissolve a little in water for the test. Notice that a red *coloration* of ferric thiocyanate, $Fe(CNS)_3$, is obtained, not a precipitate.

3. *Potassium ferricyanide* (not *ferrocyanide*). The solution of this substance must be made up fresh and, before dissolving, the outer portion of the crystal used should be removed by running water over it. Notice that there is no *precipitate,* but a dark coloration.

FIG. 25.

EXPERIMENT 120.—**Ferrous salts.**—Examine some iron wire, and if the wire contains any rust, remove it by scraping with sandpaper. Place about 0.5 gram of the wire in a small flask, add 100 c.c. water and 15 c.c. hydrochloric acid, and close the flask with a cork bearing a valve made as shown in Fig. 25. A glass tube fits tightly in the cork, and is connected with a piece of rubber tubing 4 cm. long, the end of which is closed by a bit of glass rod. In the middle of the rubber tube is a longitudinal slit 1 cm. long, made with the point

of a sharp knife-blade. This arrangement permits the escape of the hydrogen from the flask, but obstructs the entrance of air, which would convert the *ferrous* chloride into *ferric* compounds.

When the iron is dissolved, cool the solution of ferrous chloride by running water over the flask, and investigate the behavior of small portions with the same three substances employed in the preceding experiment. Keep the flask corked.

Notice that *ferrous* hydroxide, obtained by the action of ammonia, is at first white, but rapidly turns green, and finally becomes rust-colored *ferric* hydroxide. This change is due to absorption of oxygen from the air.

Potassium thiocyanate should produce no color. If a faint pink is produced, suggest a reason for it.

Potassium ferricyanide produces a deep blue *precipitate,* quite different in character from the result of its action upon the ferric solution.

Now take a little of your ferrous solution and add bromine-water, drop by drop, until the liquid is slightly reddish. Boil the liquid. Notice the color change. Does the new color resemble that of ferric chloride—the color of the ferric ion? Test it with the same three substances, using separate small portions, and see whether it is now *ferrous* or *ferric.* Write the equation for the action of chlorine on ferrous chloride.

Place a little of your ferrous chloride solution in a test-tube, add a few drops of nitric acid, and boil. Does it seem to change to ferric chloride? Apply the three tests and ascertain.

Place a little of your ferrous chloride solution in a beaker on wire gauze, and heat it gently for 15 minutes, allowing free air access. Test it for ferric compounds with potassium-thiocyanate solution. What is the action of air on ferrous salts?

EXPERIMENT 121.—**Reduction of ferric to ferrous solutions by nascent hydrogen.**—Dissolve not more than 0.1 gram of ferric chloride in 100 c.c. of water. Carefully clean the flask in which you dissolved the iron, place the ferric chloride solution in it, incline the flask, and slide in 3 grams of granulated zinc or cuttings of sheet zinc, add 10 c.c. of hydrochloric acid, insert the valve, place the flask on a piece of asbestos board, and apply a *gentle* heat. The flask should be heated at least ten minutes after the liquid has become colorless. Determine whether the solution is ferrous or ferric by the same three tests. What is the action of nascent hydrogen on ferric solutions? Write the equation. Hydrogen passed into a ferric solution through a glass tube has no effect. What is the theoretical explanation of this fact?

PROBLEMS

143. What volume of hydrogen at 13° and 780 mm. is needed to convert 31.5 grams of Fe_2O_3 into iron?

144. What volume of hydrogen at 14° and 740 mm. is required to change 20 grams of ferric oxide into iron?

145. If ferric oxide, Fe_2O_3, contains 70 per cent iron and 30 per cent oxygen, what is the atomic weight of iron?

146. If ferrous oxide, FeO, contains 77.8 per cent iron and 22.8 per cent oxygen, what is the atomic weight of iron?

CHAPTER XL

EXPERIMENT 122.—**Platinum.**—Dip some asbestos fiber into a solution of *platinic chloride,* $PtCl_4$. *Use the solution sparingly, and return any unused portion to the bottle. It is expensive.* Heat the asbestos to redness for a few seconds, holding it in forceps. The result of this operation is to coat the asbestos with finely divided plat-

inum. What gas must have escaped? What is the effect of heat upon compounds of platinum?

Let the "platinized asbestos," as it is called, cool, hold it in forceps, and allow hydrogen from a jet connected with a Kipp generator to stream out against it. The generator must be free from air, or an explosion will result. If there is any doubt about this, let hydrogen escape from it at the rate of a bubble a second, dipping the exit-tube in water for ten minutes before the experiment is tried. Describe and explain the result.

CHAPTER XLI

CARBON

EXPERIMENT 123.—**Allotropic forms of carbon.**—Examine graphite, charcoal, anthracite and bituminous coal. Describe them. Describe the diamond. Make some lampblack by holding a piece of chalk in a luminous gas flame and examine and describe it.

Color some water faintly with litmus. The liquid must be clear and free from any solid matter. Filter it, if necessary, and add to it 3–5 grams of *animal charcoal* (bone-black). Warm the liquid gently, shake for a time, and filter. If the liquid is not colorless, filter it again. Repeat, using water colored with a *little* ink (one drop) or indigo, instead of the litmus. This remarkable property of decolorizing liquids is also possessed by wood charcoal, but to a far less degree.

EXPERIMENT 124.—**Illuminating gas.**—Fill a hard glass test-tube one-quarter with bituminous coal, clamp it horizontally near the cork, and apply a heat, at first gentle and finally intense. During the heating disconnect sev-

eral times and test the gas in the hard glass tube for ammonia with red litmus paper. Notice the production of tar during the heating. Collect the gas given off over water in wide test-tubes. Examine it, especially with regard to combustibility. Examine the residue left in the tube, which is *coke.*

CHAPTER XLII

CARBON DIOXIDE AND CARBON MONOXIDE

You have already made experiments which prove that carbon dioxide is produced by the burning of charcoal in the air or in oxygen, and that it is contained in the atmosphere and in the gases from the lungs. Read again your notes of these experiments, and if you have forgotten any essential point, repeat them.

Experiment 125.—**Production of carbon dioxide by combustion.**—Fill a deflagrating spoon with powdered charcoal, heat it to redness, and let it burn in a covered bottle. Test the gas with lime-water.

Hold a dry clean bottle over a *small* gas flame for a few seconds and apply the lime-water test. Repeat with a candle flame. With burning wood (match-stick). With the flame of kerosene. Of alcohol. (The last two flames can be conveniently obtained by dipping a bunch of asbestos into the corresponding liquid, placing it on an iron plate and setting fire to it.

Draw conclusions regarding the existence of carbon in combustibles and the products of their combustion. State the evidence.

Experiment 126.—**Production of carbon dioxide by the action of acids on carbonates.**—Place a little sodium

carbonate in each of three small beakers, and set each in a larger beaker or in a bottle or cylinder. Add to the first hydrochloric acid, to the second nitric, and to the third sulphuric. Cover the larger vessels with paper. When the action is over, remove the small beakers with forceps, add lime-water to each of the larger vessels, and shake. What does this experiment prove? Write the three equations.

EXPERIMENT 127.—**Preparation of carbon dioxide.**—Use the apparatus employed in making hydrogen. Place the gas-bottle almost horizontal and slide enough broken marble into it to fill it when upright to the depth of 1 cm. If the lumps are too large they can be broken with a hammer on an anvil—not in a mortar. Add 50 c.c. of water, and then hydrochloric acid, slowly through the funnel-tube until a brisk evolution of gas is obtained. Collect it by downward displacement in dry bottles.

Investigate the physical properties of the gas. Has it any odor or taste? If in doubt about the last point, let the gas from the generator bubble through a little water in a test-tube and taste the liquid. Test its solubility in water in the usual way.

Lower a lighted candle into a bottle of the gas. In order to illustrate its high density, place a lighted candle on your desk and pour carbon dioxide over it, just as you would pour water. Balance a large beaker on the platform scales and pour a bottle of the gas into it.

A few substances, which produce very high temperatures by their combustion, will burn in the gas. Try a piece of magnesium ribbon, held in forceps and started burning in the air. What is the black substance which is obtained along with the magnesium oxide?

Pass carbon dioxide through some lime-water in a test-tube for some time. Notice that the calcium carbonate which it at first precipitated finally redissolves. Calcium

carbonate is soluble in water containing carbon dioxide, and this solution is present in many "hard waters."

Boil the liquid, and show that when the carbon dioxide is expelled the precipitate is again obtained.

EXPERIMENT 128.—**Action of glowing charcoal on carbon dioxide.**—Pass carbon dioxide from a Kipp generator over a column of glowing charcoal about 20 cm. long in a combustion-tube. The tube is heated slowly and carefully by a wing-top burner. Red rubber stoppers are best for it, but ordinary corks will answer with care. From this tube the gas passes through a solution of potassium hydroxide in a flask with a doubly perforated cork. It enters through a tube which goes almost to the bottom of the flask, and leaves through a tube which ends just inside the cork and does not dip into the liquid. This tube conveys the gas to a pan, where it is collected over water in wide test-tubes. The current of gas must be very slow.

Study the properties of the carbon monoxide, CO. Especially notice its combustion. *It is highly poisonous, and must not be inhaled.* Apply the lime-water test to the product of its combustion.

EXPERIMENT 129.—**Preparation of carbon monoxide by heating oxalic acid with sulphuric acid.**—Place 8–10 grams of oxalic acid and 50–70 grams of strong sulphuric acid in a small flask. Fit the flask with a doubly perforated cork, one hole of which carries a safety-funnel. A little mercury must be placed in the bend of the safety-funnel. Apply a very gentle heat. The gas is led through a wash-bottle containing potassium hydroxide and collected over water. The first two bottles of gas contain air, and should be allowed to escape under the hood. Investigate the properties of the gas, study its combustion by burning a bottle full of it, and test the product of the burning by lime-water.

Care must be taken not to inhale the gas produced in this experiment. In heating the flask, use a small flame, which should be removed or turned down if the contents begin to froth up. Keep the hand away from under the flask in heating, for it may break, and hot sulphuric acid attacks the flesh energetically.

Questions

1. Why is carbon monoxide poisonous?
2. Why is it no longer poisonous when the air pressure is greatly increased?
3. What volume of carbon dioxide will be produced by the burning of a liter of carbon monoxide? Why? What volume of oxygen is needed?

Problems

147. The great German chemical works, the Badische Anilin- und Soda Fabrik, burns 190,000 tons of coal a year. If the coal contains 70 per cent of carbon, and if there are 310 working days in the year, what weight of CO_2 escapes daily from the chimneys of the establishment?

148. What volume of carbon dioxide is formed by the burning of 30 liters of carbon monoxide, and what volume of oxygen is required? Solve by inspection.

149. What gas gives rise to the blue flame often seen playing over the surface of a coal fire? How much coal containing 90 per cent of carbon would be needed to make 5,000 liters of this gas at 15° and 750 mm.?

150. What volume of gas measured under standard conditions is produced when 20 grams of pure dry oxalic acid are heated with strong sulphuric acid? What is the composition of the gas?

151. What volume of carbon dioxide must be passed over glowing charcoal to form 42 grams of carbon monoxide?

152. What is the volume (*a*) of 50 grams of carbon monoxide? (*b*) Of 50 grams of carbon dioxide?

153. What volume of carbon dioxide would be produced by burning a diamond weighing 3 grams in oxygen?

154. How much carbon is there in (*a*) 2.8 liters of carbon dioxide? (*b*) 2.8 liters of carbon monoxide?

155. What volume of carbon dioxide at standard conditions is produced by dissolving 25 grams of marble in hydrochloric acid?

156. What volume of carbon dioxide at 12° and 750 mm. is produced by dissolving 20 grams of marble in hydrochloric acid?

CHAPTER XLIII

SOME CARBON COMPOUNDS

EXPERIMENT 130.—**Methane.**—Heat some *sodium acetate* in an iron dish until the water of crystallization has escaped and the melted dry salt is left. Mix 3 grams of this in a mortar with 3 grams of lime and 3 grams of sodium hydroxide. Grind thoroughly and transfer the mixture to a hard glass test-tube. Clamp the tube horizontal and fit it with a cork bearing a delivery tube. Apply a gentle heat, gradually increased to redness. Collect the gas in wide test-tubes. Cover the hard glass test-tube with soot, and disconnect before letting it cool.

Record the physical properties of the gas. Burn some, notice the character of the combustion, and apply the lime-water test.

EXPERIMENT 131.—**Acetylene.**—Fill a small test-tube one-third with water and throw a piece of *calcium carbide* into it. Has the gas any color or odor? Light it. Does its combustion resemble that of methane? Record the physical properties of calcium carbide.

EXPERIMENT 132.—**Various carbon compounds.**—*Ether, wood alcohol, and carbon disulphide are highly inflammable, and must not be used when a flame is anywhere in the neighborhood. The bottles containing them must be kept tightly corked.*

Examine the following important carbon compounds and record their properties. Study also their solubility

in water. *Alcohol, ether, chloroform, wood alcohol* (methyl alcohol), and *carbon disulphide.* Briefly state in your notes the source and uses of each.

EXPERIMENT 133. — **Fermentation.** — Dissolve 100 grams of grape-sugar (glucose) in a liter of water. Mix half of a compressed yeast cake to a thin paste with water and add it to the liquid. Place the mixture in a flask closed by a singly perforated cork bearing a tube which leads to a small flask containing 50 c.c. of lime-water. The tube should dip into the lime-water. This flask is closed by a doubly perforated cork, and a tube from the second hole is connected with a U-tube containing pieces of sodium hydroxide. This is to prevent the CO_2 of the air from entering the flask containing the lime-water.

Let the apparatus stand several days, if possible. What change occurs in the lime-water? What gas must have been given off? What is the other product?

To obtain it, distill the liquid in the large flask in the apparatus employed for the distillation of water (Fig. 5). Continue distilling until $\frac{1}{10}$ of the liquid has passed over, and discard the rest. If there is too much liquid to be distilled in one operation, divide it into several portions, and distill $\frac{1}{10}$ of each. The flask should not be more than $\frac{2}{3}$ filled. Unite these distillates in a small flask, and distill $\frac{1}{10}$ of the liquid, using a dry test-tube as a receiver. Has the liquid which distills the odor of alcohol? Is it combustible? If not, place some in a dish, heat it, and hold a lighted match near the surface of the liquid.[1]

EXPERIMENT 134.—**Aldehyde.**—Make a strong solution of *potassium dichromate,* place it in a test-tube, and add 1 c.c. of strong sulphuric acid and 1 c.c. of alcohol. Heat the liquid. The odor produced is that of *aldehyde.* Notice the change from the red color of potassium di-

[1] This distillation consumes much time and is not always successful. It is sufficient to notice the production of CO_2 and the general nature of the process.

chromate to the green color of *chromium sulphate,* $Cr_2(SO_4)_3$.

EXPERIMENT 135.—**Acetic acid.**—Examine acetic acid and record its properties. Dilute some with 100 times its volume of water, and test the liquid with litmus paper (red and blue). Taste it. Mix 5 c.c. of strong acetic acid with 10 c.c. of alcohol. If possible, preserve some of the liquid in a corked test-tube, and notice the gradual development of the fragrant odor of *acetic ether.* This is an interesting example of slow chemical change. Sulphuric acid catalytically accelerates the process. Add to some of the fresh mixture in a test-tube 2 c.c. of strong sulphuric acid and heat gently. The odor appears at once. This is a *test* for acetic acid or for alcohol.

Examine *vinegar,* taste it, and try its behavior with red and blue litmus paper.

Examine *sodium acetate,* and notice its similarity in general character to ordinary sodium salts, like the nitrate and the sulphate.

QUESTIONS

1. What is fire-damp, and how is it formed?

2. What is the valence of carbon in CH_4, C_2H_6, C_3H_8, $CHCl_3$, C_2H_5Cl?

3. What is a *substitution product?* Give examples. Why are substitution products so numerous?

4. Write the formulas of the first nine of the series of hydrocarbons which begins with *methane.* How do the physical properties of the substances change as we advance in the series?

5. Why is it that a mixture of methane or any other combustible gas with air will *explode,* although the gas alone will burn quietly?

6. What is fermentation? State briefly two methods by which it has been proved that fermentation is not caused by any vital activity of yeast.

Problems

157. What is the weight of 38 liters of methane at 31°?

158. What is the weight of 10 liters of acetylene?

159. How much calcium carbide is needed to produce 5.6 liters of acetylene?

160. A town is to be lighted with acetylene. It is calculated that the consumption of the gas will be 70,000 liters per day. How much calcium carbide will be required per month of thirty days?

161. How much carbon is there in 32 liters of acetylene?

162. How much sodium acetate is needed to make 8 liters of methane?

$$NaC_2H_3O_2 + NaOH = CH_4 + Na_2CO_3.$$

163. How many liters of oxygen at 10° and 780 mm. are needed to burn completely the methane obtained when 41 grams of sodium acetate are heated with sodium hydroxide?

$$NaC_2H_3O_2 + NaOH = CH_4 + Na_2CO_3.$$
$$CH_4 + 2O_2 = CO_2 + 2H_2O.$$

164. Calcium carbide is treated with water and the acetylene burned. 28 liters of carbon dioxide at 15° and 740 mm. resulted from the combustion. How much calcium carbide was taken, and what volume of oxygen at 15° and 740 mm. was needed to burn the acetylene?

$$CaC_2 + 2H_2O = Ca(OH)_2 + C_2H_2.$$
$$C_2H_2 + 5O = 2CO_2 + H_2O.$$

CHAPTER XLIV

ADDITIONAL CARBON COMPOUNDS

Experiment 136.—**Soap.**—Dissolve 25 grams of sodium hydroxide in 150 c.c. of water in an iron dish. Add 75 grams of *lard,* and boil gently for half an hour or more. Slowly add 50 grams of salt, stirring constantly. *Be careful during the boiling not to allow any of the*

liquid to be spattered into the eyes. It is well to keep a glass plate between the dish and the face.

The solid which separates at the top is *soap.* What else has been produced, and where is it? Remove the soap and examine it. It is a mixture of the sodium salts of *palmitic, stearic,* and *oleic acids.* Dissolve about 3 grams of shavings of it in 100 c.c. of warm water in a dish. Add dilute sulphuric acid to a portion of it. The mass which separates consists of a mixture of the acids mentioned above, which are liberated from their sodium salts by the sulphuric acid, sodium sulphate being formed. To another small portion of the soap solution add a dilute solution of *calcium chloride.* Water containing calcium compounds is said to be *hard.* Why can not hard water be used satisfactorily in washing with soap?

EXPERIMENT 137.—**Albumin.**—Break an egg and separate some of the white from the yolk. Dilute about 5 c.c. of the white of egg with 100 c.c. of water. Boil a portion of this liquid in a test-tube. (?) To another portion in a test-tube add a little nitric acid. The albumin is precipitated in both experiments.

Heat a bit of meat, a feather, a few clippings of horn, or a cochineal insect in a dry test-tube. Almost any form of animal matter will answer. Describe the nature of the change. Notice the odor. Show by red litmus paper that ammonia is given off. This proves that nitrogen and hydrogen are two of the constituents of albumin. Look for evidence of the presence of carbon in the residue in the tube.

CHAPTER XLV

No experiments.

CHAPTER XLVI

No experiments.

Additional exercise. The fundamental principles of qualitative analysis.[1]—A. Prepare dilute solutions of the nitrates of *silver, copper, iron, barium,* and *sodium.* About 100 c.c. of each will be needed. Copper nitrate can be made by heating a little copper in a test-tube with *dilute* nitric acid until the action ceases. The liquid should still contain undissolved copper. It is diluted and filtered if necessary, or simply poured off from the copper. Iron nitrate (ferric) can be made in a similar way, but the iron must be well covered with water and the nitric acid added drop by drop, for the action is violent. The other solutions can be made by dissolving 1 gram of the corresponding nitrate in 100 c.c. of water.

Study the behavior of small portions of each of these liquids with the following substances separately:

B. Hydrochloric acid, sulphuric acid, hydrogen sulphide, ammonium hydroxide.

The experiments should be made in test-tubes. The liquids in B should be added one drop at a time, stirring constantly. Note especially whether a precipitate or a color is produced, and if the former, whether it is soluble in excess. The hydrogen sulphide can be taken from a Kipp generator. Use a *gentle* current of the gas, not more than a bubble every two seconds. *The tube through which the H_2S passes into the liquid must be cleaned*

[1] The object of this exercise is not to teach qualitative analysis, but to give the student a grasp of the principles upon which the separation of one metal from another depends, so that he will be able afterward to attack the subject understandingly. The scheme is that of Professor Richards (Harvard Requirements, p. 20).

both from liquid and solid before using it again for a fresh test, and every test-tube and beaker employed must be absolutely clean. If no Kipp generator is available, make the hydrogen sulphide yourself by the method already studied. *Be careful not to inhale it unnecessarily.*

NOTE.—The precipitate formed when the solution of ferric nitrate is treated with H_2S contains no iron. It is sulphur. The change in the color of the liquid will show you that the iron is now in the *ferrous* condition.

QUESTIONS

1. Devise a method of separating *silver* from any of the metals in A based on the use of hydrochloric acid. Try it, mixing a little of the silver-nitrate and copper-nitrate solutions for the purpose. Why is it necessary to wash the precipitate by filling the filter with water and letting it drain off half a dozen times after filtering? This must *always* be done in work of this kind. In the present exercise the wash liquid can be thrown away. Of course, if the work is *quantitative* it must be retained. Why is it necessary to add the hydrochloric acid until it produces no further precipitate? This is *always* necessary in similar cases. How can you prove that the precipitate is really silver chloride? How could you tell after filtering whether the silver had all been precipitated or not?

2. Devise a method of separating *barium* from any of the others. Try it with barium and copper nitrates.

3. Devise and carry out a method of separating silver, copper, and barium, when all three are present together.

4. Devise and carry out a method of separating copper from sodium by hydrogen sulphide. Evaporate the liquid, which should contain the sodium, to dryness, and apply the flame test. Identify the copper as in 5 *b* (next paragraph).

5. Separate copper from iron by hydrogen sulphide. (*a*) Heat the filtrate containing the *ferrous* ions to gentle boiling for twenty minutes to drive off the H_2S. Add a few drops of nitric acid, and boil. What change occurs in the condition of the iron? Show that the liquid contains iron, using ammonium hydroxide. (*b*) Dissolve the copper sulphide by repeatedly pouring the same portion of hot dilute nitric acid through the filter. The blue liquid

contains *copper nitrate*. Refer to your previous notes for several methods of identifying copper in it, and apply them.

6. Separate iron from sodium, using ammonium hydroxide.

7. Devise and carry out a method of separating copper, iron, and sodium when all three are present together.

8. Devise and carry out a method for the separation of all five of the metals in A, using only the substances in B for the separation. Any method you are familiar with can be used to identify the metals when separated.

APPENDIX

CALCULATION OF THE EFFECT OF TEMPERATURE, PRESSURE, AND WATER-VAPOR ON THE VOLUMES OF GASES

1. **Temperature.**—The *absolute temperature* is the temperature measured from 273° below 0° C. Thus, the absolute temperature of 10° is $273 + 10 = 283°$.

The absolute temperature of —10° is $273 - 10 = 263°$.

PROBLEM

165. Calculate the absolute temperature corresponding to the following centigrade temperatures:

a. 13°. *b.* 274°. *c.* −50°. *d.* −278°.

The volume of a mass of gas is directly proportional to its absolute temperature. Let T and t = two temperatures—both absolute. Let V_T be the volume at $T°$ and V be the volume at $t°$. Then—

$$V_T = V_t \times \frac{T}{t}.$$

In words, this means if you know the volume of a mass of gas at some known temperature, you can calculate its volume at some other temperature by multiplying the old volume by the new temperature, and dividing by the old temperature, both temperatures being absolute.

Never make the error of using ordinary centigrade temperatures instead of absolute temperatures. It is easy to see that this leads directly to absurd results. For in-

stance: I have a liter of gas at 0° C. What will its volume become at 273° C.?

Here, if we use ordinary centigrade degrees, the volume becomes

$$1 \times \frac{273}{0};$$

that is, the volume is infinite—which is absurd. But if we employ absolute degrees, the volume is—

$$1 \times \frac{273}{546} = 2 \text{ liters};$$

which is the correct result.

Problems

166. What volume will a liter of air at 0° C. occupy at 100° C.?

167. 5 liters of oxygen at 0° C. occupy what volume (*a*) at 10° C.? (*b*) at −10° C.?

168. 25 c.c. of nitrogen at 15° C. will measure what at the standard temperature 0° C.?

169. I have 500 c.c. of hydrogen at 13° C. What will the volume become at 65° C.?

170. 600 c.c. of oxygen at 28° C. will measure what at −14° C.?

171. 500 liters of air at 20° C. will occupy what volume at 80° C.?

172. A liter of steam at 100° C. will occupy what volume at 120° C.?

173. 67 liters of air are heated from −30° C. to 60° C.? What is the new volume?

Since the volume of a mass of gas varies with the temperatures, it is always necessary, in measuring gases, to know the temperature of the gas measured. And it is clear that the expression "1 liter of oxygen" has no meaning unless some particular temperature is either stated or understood. Now, in order to avoid the necessity of continually stating the temperature, it is extremely convenient to assume some temperature as a standard

point which is to be understood unless some other temperature is stated. *The standard temperature universally agreed upon is 0° C.—the melting-point of ice.* Thus, when a writer speaks of "1 liter of oxygen" without stating the temperature under which the gas was measured, we know that 0° C. is meant.

The student should grasp the fact that every problem like those just solved is supposed to deal with a certain fixed weight of gas which is not added to or subtracted from during the process of heating or cooling. Clearly, if temperature and pressure remain the same, the volume must be directly proportional to the weight of the gas. Thus, 1 gram of hydrogen at standard temperature and pressure occupies a volume of 11.2 liters. Evidently, 2 grams of hydrogen must measure 22.4 liters under the same conditions, and so on. But, in all problems of this sort, the quantity of gas is supposed to remain the same.

2. **Pressure.**—*The volume of a mass of gas is inversely proportional to the pressure upon it.* If p and P are two pressure, both stated in millimeters of mercury, and if V_p and V_P are the volumes which the same quantity of gas will occupy at those pressures, then

$$V_P : V_p :: p : P. \text{ Hence}$$

$$V_P = V_p \times \frac{p}{P}.$$

Therefore, if the volume of a mass of gas is given at some pressure and it is required to calculate its volume at some other pressure, we must multiply the old volume by the old pressure and divide by the new pressure.

Problems

174. 10 liters of gas at a pressure of 743 mm. will occupy what volume at 720 mm.?

175. 18.5 c.c. of nitrogen are measured under a pressure of 745 mm. What will the volume be at 760 mm.?

176. A liter of oxygen is measured at 760 mm. What will it measure at 748 mm. ?

177. 100 c.c. of air at 760 mm. (1 atmosphere) will occupy what volume under 20 atmospheres ?

178. What pressure is required to compress 500 c.c. of carbon dioxide at 728 mm. to a volume of 400 c.c. ?

179. What must the pressure be made in order to allow the 500 c.c. of gas of the preceding problem to expand to 850 c.c. ?

In order not to be compelled to state continually the pressure, in speaking of the volumes of gases, and in order to be able to compare gas volumes, measured at different temperatures, with each other, *760 mm. of mercury is agreed upon as the standard pressure, which is understood when no pressure is stated.* This pressure is called 1 atmosphere, because the pressure of the air does not vary widely from that amount.

Since, as we have seen, 0° is the standard temperature, the expression "standard conditions" means 0° and 760 mm. Thus, when a writer speaks of 1 liter of oxygen (or of any volume of any gas) without mentioning either temperature or pressure, we understand at once that the gas is supposed to exist at 0°, and under a pressure of 760 mm.

3. *When temperature and pressure both vary,* we have simply to correct for both by the methods already studied. This can be done in two separate calculations, but it is easier and better to unite both corrections in one operation. The volume of a gas is directly proportional to the absolute temperature and inversely proportional to the pressure. Let

V_{PT} = the volume at the absolute temperature T and pressure P.

V_{pt} = the volume at the absolute temperature t and pressure p.

Then
$$V_{PT} = V_{pt} \times \frac{T}{t} \times \frac{p}{P}.$$

In words, this means that in order to calculate the new volume of a gas at some new temperature and pressure, *we must multiply the old volume by the new temperature and the old pressure, and divide it by the old temperature and the new pressure.* Of course, both temperatures must be absolute.

Such calculations can be rapidly, easily, and correctly made by the use of *logarithms,* and this is true of chemical calculations generally. A table of logarithms is given for this purpose, and its use will save about half the time and labor of chemical calculation, and will greatly reduce the number of errors in the numerical work.

Problems

180. 100 c.c. of oxygen at 15° C. and 740 mm. will occupy what volume at standard conditions?

$$V_{PT} = 100 \times \frac{273}{288} \times \frac{740}{760} = 92.3 \text{ c.c.}$$

The student will find that his chief difficulty in solving problems like this and the following ones is in determining which temperature and pressure to put in the numerator and which in the denominator. It will pay to make it a rule to inspect the fractions with great care before working out the calculation. Errors can be detected by the exercise of a little common sense. For instance, in the preceding problem the gas is to be cooled from 15° C. to 0° C. This will *reduce* its volume. Hence, the temperature-fraction must be $\frac{273}{288}$, not $\frac{288}{273}$. Also, the pressure is to be raised from 740 to 760, and this also will *reduce* the volume. Hence, the pressure fraction must be $\frac{740}{760}$, not $\frac{760}{740}$.

181. Supposing the initial temperature in the preceding problem to be −15° C. instead of 15° C., what would be the new volume? The other figures remain the same.

182. What volume will 48 c.c. of nitrogen at standard conditions occupy at 18° C. and 733 mm. ?

183. 25 liters of a gas at standard conditions are cooled to −10° C., and the pressure reduced to 723 mm. What is the new volume ?

184. 310 c.c. of hydrogen at 10° C. and 530 mm. will occupy what volume at 18.7° C. and 590 mm. ?

185. 1,704 c.c. of nitrogen at 11° C. and 760 mm. are brought to a temperature of 27° C. and a pressure of 900 mm. What is the volume ?

186. 271 c.c. of hydrogen at 269° and 900 mm. are cooled to −51° C., and the pressure decreased to 666 mm. Calculate the final volume.

4. *The effect of water-vapor on the volume of a mass of gas.*—Suppose that we have 100 c.c. of dry oxygen confined over mercury in a graduated tube. Let us admit a drop of water and allow the oxygen to saturate itself with moisture. Clearly, the volume of gas in the tube must increase, for the water-vapor will occupy space. The result is the same as though we had introduced a little nitrogen or some other gas into the tube and allowed it to mix with the oxygen.

The volume can be kept 100 c.c. by increasing the pressure under which the gas is measured. But if this is done, the total pressure can not be considered as exerted upon the oxygen in the tube, for the water-vapor is also present. Hence, the pressure under which the gas really exists and is measured is *less* than the total pressure. How much less?

The pressure which saturated water-vapor exerts at various temperatures is given in the table. When a gas is measured *over water,* or when it is measured saturated with water, *the pressure which water-vapor exerts at the temperature of measurement must be ascertained from the table and deducted from the total pressure. The remainder will be the pressure under which the gas is really measured.*

Problems

187. A mass of air at 15.3° C. and 747.2 mm., measured *over water*, occupied a volume of 82.4 c.c. What volume would it occupy dry and at standard conditions?

Solution

From the table we observe that water-vapor at 15° C. exerts a pressure of 12.7 mm. and at 16° C. a pressure of 13.54 mm. Hence its pressure at 15.3° C. must = 12.9 mm.

The pressure under which the gas is really measured is

$$747.2 - 12.9 = 734.3 \text{ mm.}$$

The rest of the calculation is the same as in the preceding problems:

$$82.4 \times \frac{273}{288.3} \times \frac{734.3}{760} = 75.39 \text{ c.c.}$$

188. 11.41 c.c. of a mixture of oxygen and hydrogen are measured over water at 14° C. and 743 mm. Calculate the volume under standard conditions.

189. 112.1 c.c. of nitrogen saturated with water at 16° C. and 744 mm. will occupy what volume dry and under standard conditions?

190. The gas-holder of a gas-works contains 4,500 cubic meters of illuminating gas, confined over water. The temperature is 9° C. and the pressure 776 mm. How many cubic meters would the gas measure under standard conditions?

191. 100 c.c. of oxygen are confined over water and measured at 14° C. and 756 mm. What will be the volume when the gas is dried and placed under standard conditions?

192. A gas-holder contains 10 liters of air confined over water at 20° C. and 756 mm. What will the gas measure when dried, other conditions remaining the same?

Table of Atomic Weights

	Exact values.	Approximate values.		Exact values.	Approximate values.
Aluminium ...	27.1	27	Neodymium ..	143.6	....
Antimony.....	120.0	120	Neon.....	19.94	....
Argon	39.92	....	Nickel	58.70	58.7
Arsenic	75.0	75	Nitrogen	14.04	14
Barium	137.43	137	Osmium.....	190.8	....
Bismuth.....	208.0	208	Oxygen	16.000	16
Boron	11.0	11	Palladium....	106.5	....
Bromine	79.955	80	Phosphorus...	31.0	31
Cadmium	112.3	112	Platinum.....	195.2	195
Caesium.....	132.9	....	Potassium	39.14	39
Calcium.....	40.1	40	Praseodymium	140.5	....
Carbon.....	12.001	12	Rhodium.....	103.0	....
Cerium.....	140.0	....	Rubidium	85.44	....
Chlorine	35.455	35.5	Ruthenium...	101.7	....
Chromium	52.14	52	Samarium	150.0	....
Cobalt	59.00	59	Scandium.....	44.0	....
Columbium....	94.0	....	Selenium.....	79.2	....
Copper.....	63.60	63.5	Silicon	28.4	28.5
Erbium	166.0	....	Silver	107.93	108
Fluorine	19.05	19	Sodium	23.05	23
Gadolinium ...	156. ?	....	Strontium	87.68	87.5
Gallium.....	70.0	....	Sulphur	32.065	32
Germanium ...	72.5	....	Tantalum	183.0	....
Glucinum.....	9.1	....	Tellurium	127.5 ?	....
Gold.....	197.3	197	Terbium	160.0	....
Helium	3.96	....	Thallium.....	204.15	....
Hydrogen.....	1.0075	1	Thorium	233.0	....
Indium	114.0	....	Thulium.....	171.0 ?	....
Iodine	126.85	127	Tin	119.0	119
Iridium.....	193.0	....	Titanium.....	48.17	....
Iron	55.9	56	Tungsten.....	184.0	....
Krypton	81.7	....	Uranium	238.5	239
Lanthanum ...	138.5	....	Vanadium.....	51.4	....
Lead	206.92	207	Xenon	128.0	....
Lithium.....	7.03	....	Ytterbium....	173.0	....
Magnesium ...	24.36	24	Yttrium.....	89.0	....
Manganese....	55.02	55	Zinc.....	65.40	65.5
Mercury.....	200.0	....	Zirconium....	90.6	....
Molybdenum ..	96.0	....			

This table contains the values of the atomic weights according to the calculations of Prof. Richards. The *approximate* values should be used in solving all problems.

Vapor Pressure of Water

Temperature, Centigrade.	Vapor pressure in mm. of mercury.	Temperature, Centigrade.	Vapor pressure in mm. of mercury.	Temperature, Centigrade.	Vapor pressure in mm. of mercury.
−10	2.09	12	10.46	26	24.99
− 5	3.11	13	11.16	27	26.51
0	4.60	14	11.91	28	28.10
+ 1	4.94	15	12.70	29	29.78
2	5.30	16	13.54	30	31.55
3	5.69	17	14.42	35	41.83
4	6.10	18	15.36	40	54.91
5	6.53	19	16.35	50	91.98
6	7.00	20	17.39	60	148.79
7	7.49	21	18.50	70	233.09
8	8.02	22	19.66	80	354.64
9	8.57	23	20.89	90	525.45
10	9.17	24	22.18	100	760.00
11	9.79	25	23.55		

The vapor pressure of water for a temperature not given in the table can easily be found by calculation. Thus, suppose it is required to find the vapor pressure for the temperature of 32.5°. The increase in vapor pressure from 30° to 35° is $41.83 - 31.55 = 10.28$ mm. Hence, the increase from 30° to 32.5° will not be far from $10.28 \times \frac{2.5}{5}$, or 5.14 mm., and the vapor pressure for 32.5° will be about 36.69 mm. It will not be exactly 36.69 mm., because, in the calculation, it is assumed that the vapor pressure increases proportionally with the temperature, which is not the case, but for small differences of temperature the error is small.

When a gas is measured over water, or moist, the vapor pressure of water for the temperature must be subtracted from the pressure under which the gas is measured.

Factors for converting Metric into Ordinary Units

1 inch = 2.54 cm. 1 cm. = 0.3937 in.

For practical purposes it is sufficient to remember that about $2\frac{1}{2}$ cm. = 1 in.

A *liter* is the volume of a cube whose side is 10 cm. Therefore,

1 liter = 1,000 cubic centimeters;
1 pint = 0.5679 liter;
1 gallon = 4.54346 liters;
1 liter = 0.2201 gallon.

The *gram* is the weight of 1 c.c. of pure water at 4°.

1 liter of pure water at 4° = 1 kilo (1,000 grams).

1 oz. = 28.35 gms. 1 gram = 15.432 grains.
1 lb. = 453.6 gms. 1 kilo = 2.2046 lbs.

Formulæ for converting Fahrenheit Degrees into Centigrade, and the Reverse

$$C.^\circ = \frac{5}{9}(F.^\circ - 32).$$

$$F.^\circ = \frac{9}{5}C.^\circ + 32.$$

Miscellaneous Data

Weight, 1 liter pure hydrogen at 0° and 760 mm. = .0896 gram. Volume of molecular weight in grams of any gas or vapor, 22.4 liters.

Melting-point of	ice	0° C.
"	" alcohol	−130° C.
"	" mercury	−39.4° C.
"	" hydrogen, about	−260° C.
"	" tin	233° C.
"	" lead	334° C.
"	" silver	954° C.
"	" gold	1,064° C.
"	" iron (pure)	1,600° C.

Boiling-point of	water	100° C.
"	" alcohol	78° C.
"	" ether	35° C.
"	" chloroform	61.5° C.
"	" carbon disulphide	47° C.
"	" hydrogen	−252.8° C.
"	" oxygen	−183° C.

Air contains 21 per cent (20.97 per cent) by volume of oxygen. 1 liter of dry air at 0° and 760 mm. weighs 1.293 grams.

Nat. Numbers	0	1	2	3	4	5	6	7	8	9	Proportional parts 1 2 3	4 5 6	7 8 9
10	0000	0043	0086	0128	0170	0212	0253	0294	0334	0374	4 8 12	17 21 25	29 33 37
11	0414	0453	0492	0531	0569	0607	0645	0682	0719	0755	4 8 11	15 19 23	26 30 34
12	0792	0828	0864	0899	0934	0969	1004	1038	1072	1106	3 7 10	14 17 21	24 28 31
13	1139	1173	1206	1239	1271	1303	1335	1367	1399	1430	3 6 10	13 16 19	23 26 29
14	1461	1492	1523	1553	1584	1614	1644	1673	1703	1732	3 6 9	12 15 18	21 24 27
15	1761	1790	1818	1847	1875	1903	1931	1959	1987	2014	3 6 8	11 14 17	20 22 25
16	2041	2068	2095	2122	2148	2175	2201	2227	2253	2279	3 5 8	11 13 16	18 21 24
17	2304	2330	2355	2380	2405	2430	2455	2480	2504	2529	2 5 7	10 12 15	17 20 22
18	2553	2577	2601	2625	2648	2672	2695	2718	2742	2765	2 5 7	9 12 14	16 19 21
19	2788	2810	2833	2856	2878	2900	2923	2945	2967	2989	2 4 7	9 11 13	16 18 20
20	3010	3032	3054	3075	3096	3118	3139	3160	3181	3201	2 4 6	8 11 13	15 17 19
21	3222	3243	3263	3284	3304	3324	3345	3365	3385	3404	2 4 6	8 10 12	14 16 18
22	3424	3444	3464	3483	3502	3522	3541	3560	3579	3598	2 4 6	8 10 12	14 15 17
23	3617	3636	3655	3674	3692	3711	3729	3747	3766	3784	2 4 6	7 9 11	13 15 17
24	3802	3820	3838	3856	3874	3892	3909	3927	3945	3962	2 4 5	7 9 11	12 14 16
25	3979	3997	4014	4031	4048	4065	4082	4099	4116	4133	2 3 5	7 9 10	12 14 15
26	4150	4166	4183	4200	4216	4232	4249	4265	4281	4298	2 3 5	7 8 10	11 13 15
27	4314	4330	4346	4362	4378	4393	4409	4425	4440	4456	2 3 5	6 8 9	11 13 14
28	4472	4487	4502	4518	4533	4548	4564	4579	4594	4609	2 3 5	6 8 9	11 12 14
29	4624	4639	4654	4669	4683	4698	4713	4728	4742	4757	1 3 4	6 7 9	10 12 13
30	4771	4786	4800	4814	4829	4843	4857	4871	4886	4900	1 3 4	6 7 9	10 11 13
31	4914	4928	4942	4955	4969	4983	4997	5011	5024	5038	1 3 4	6 7 8	10 11 12
32	5051	5065	5079	5092	5105	5119	5132	5145	5159	5172	1 3 4	5 7 8	9 11 12
33	5185	5198	5211	5224	5237	5250	5263	5276	5289	5302	1 3 4	5 6 8	9 10 12
34	5315	5328	5340	5353	5366	5378	5391	5403	5416	5428	1 3 4	5 6 8	9 10 11
35	5441	5453	5465	5478	5490	5502	5514	5527	5539	5551	1 2 4	5 6 7	9 10 11
36	5563	5575	5587	5599	5611	5623	5635	5647	5658	5670	1 2 4	5 6 7	8 10 11
37	5682	5694	5705	5717	5729	5740	5752	5763	5775	5786	1 2 3	5 6 7	8 9 10
38	5798	5809	5821	5832	5843	5855	5866	5877	5888	5899	1 2 3	5 6 7	8 9 10
39	5911	5922	5933	5944	5955	5966	5977	5988	5999	6010	1 2 3	4 5 7	8 9 10
40	6021	6031	6042	6053	6064	6075	6085	6096	6107	6117	1 2 3	4 5 6	8 9 10
41	6128	6138	6149	6160	6170	6180	6191	6201	6212	6222	1 2 3	4 5 6	7 8 9
42	6232	6243	6253	6263	6274	6284	6294	6304	6314	6325	1 2 3	4 5 6	7 8 9
43	6335	6345	6355	6365	6375	6385	6395	6405	6415	6425	1 2 3	4 5 6	7 8 9
44	6435	6444	6454	6464	6474	6484	6493	6503	6513	6522	1 2 3	4 5 6	7 8 9
45	6532	6542	6551	6561	6571	6580	6590	6599	6609	6618	1 2 3	4 5 6	7 8 9
46	6628	6637	6646	6656	6665	6675	6684	6693	6702	6712	1 2 3	4 5 6	7 7 8
47	6721	6730	6739	6749	6758	6767	6776	6785	6794	6803	1 2 3	4 5 5	6 7 8
48	6812	6821	6830	6839	6848	6857	6866	6875	6884	6893	1 2 3	4 4 5	6 7 8
49	6902	6911	6920	6928	6937	6946	6955	6964	6972	6981	1 2 3	4 4 5	6 7 8
50	6990	6998	7007	7016	7024	7033	7042	7050	7059	7067	1 2 3	3 4 5	6 7 8
51	7076	7084	7093	7101	7110	7118	7126	7135	7143	7152	1 2 3	3 4 5	6 7 8
52	7160	7168	7177	7185	7193	7202	7210	7218	7226	7235	1 2 2	3 4 5	6 7 7
53	7243	7251	7259	7267	7275	7284	7292	7300	7308	7316	1 2 2	3 4 5	6 6 7
54	7324	7332	7340	7348	7356	7364	7372	7380	7388	7396	1 2 2	3 4 5	6 6 7
	0	1	2	3	4	5	6	7	8	9	1 2 3	4 5 6	7 8 9

Nat. Number	0	1	2	3	4	5	6	7	8	9	Proportional parts 1	2	3	4	5	6	7	8	9
55	7404	7412	7419	7427	7435	7443	7451	7459	7466	7474	1	2	2	3	4	5	5	6	7
56	7482	7490	7497	7505	7513	7520	7528	7536	7543	7551	1	2	2	3	4	5	5	6	7
57	7559	7566	7574	7582	7589	7597	7604	7612	7619	7627	1	2	2	3	4	5	5	6	7
58	7634	7642	7649	7657	7664	7672	7679	7686	7694	7701	1	1	2	3	4	4	5	6	7
59	7709	7716	7723	7731	7738	7745	7752	7760	7767	7774	1	1	2	3	4	4	5	6	7
60	7782	7789	7796	7803	7810	7818	7825	7832	7839	7846	1	1	2	3	4	4	5	6	6
61	7853	7860	7868	7875	7882	7889	7896	7903	7910	7917	1	1	2	3	4	4	5	6	6
62	7924	7931	7938	7945	7952	7959	7966	7973	7980	7987	1	1	2	3	3	4	5	6	6
63	7993	8000	8007	8014	8021	8028	8035	8041	8048	8055	1	1	2	3	3	4	5	5	6
64	8062	8069	8075	8082	8089	8096	8102	8109	8116	8122	1	1	2	3	3	4	5	5	6
65	8129	8136	8142	8149	8156	8162	8169	8176	8182	8189	1	1	2	3	3	4	5	5	6
66	8195	8202	8209	8215	8222	8228	8235	8241	8248	8254	1	1	2	3	3	4	5	5	6
67	8261	8267	8274	8280	8287	8293	8299	8306	8312	8319	1	1	2	3	3	4	5	5	6
68	8325	8331	8338	8344	8351	8357	8363	8370	8376	8382	1	1	2	3	3	4	4	5	6
69	8388	8395	8401	8407	8414	8420	8426	8432	8439	8445	1	1	2	2	3	4	4	5	6
70	8451	8457	8463	8470	8476	8482	8488	8494	8500	8506	1	1	2	2	3	4	4	5	6
71	8513	8519	8525	8531	8537	8543	8549	8555	8561	8567	1	1	2	2	3	4	4	5	5
72	8573	8579	8585	8591	8597	8603	8609	8615	8621	8627	1	1	2	2	3	4	4	5	5
73	8633	8639	8645	8651	8657	8663	8669	8675	8681	8686	1	1	2	2	3	4	4	5	5
74	8692	8698	8704	8710	8716	8722	8727	8733	8739	8745	1	1	2	2	3	4	4	5	5
75	8751	8756	8762	8768	8774	8779	8785	8791	8797	8802	1	1	2	2	3	3	4	5	5
76	8808	8814	8820	8825	8831	8837	8842	8848	8854	8859	1	1	2	2	3	3	4	5	5
77	8865	8871	8876	8882	8887	8893	8899	8904	8910	8915	1	1	2	2	3	3	4	4	5
78	8921	8927	8932	8938	8943	8949	8954	8960	8965	8971	1	1	2	2	3	3	4	4	5
79	8976	8982	8987	8993	8998	9004	9009	9015	9020	9025	1	1	2	2	3	3	4	4	5
80	9031	9036	9042	9047	9053	9058	9063	9069	9074	9079	1	1	2	2	3	3	4	4	5
81	9085	9090	9096	9101	9106	9112	9117	9122	9128	9133	1	1	2	2	3	3	4	4	5
82	9138	9143	9149	9154	9159	9165	9170	9175	9180	9186	1	1	2	2	3	3	4	4	5
83	9191	9196	9201	9206	9212	9217	9222	9227	9232	9238	1	1	2	2	3	3	4	4	5
84	9243	9248	9253	9258	9263	9269	9274	9279	9284	9289	1	1	2	2	3	3	4	4	5
85	9294	9299	9304	9309	9315	9320	9325	9330	9335	9340	1	1	2	2	3	3	4	4	5
86	9345	9350	9355	9360	9365	9370	9375	9380	9335	9390	1	1	2	2	3	3	4	4	5
87	9395	9400	9405	9410	9415	9420	9425	9430	9435	9440	0	1	1	2	2	3	3	4	4
88	9445	9450	9455	9460	9465	9469	9474	9479	9484	9489	0	1	1	2	2	3	3	4	4
89	9494	9499	9504	9509	9513	9518	9523	9528	9533	9538	0	1	1	2	2	3	3	4	4
90	9542	9547	9552	9557	9562	9566	9571	9576	9581	9586	0	1	1	2	2	3	3	4	4
91	9590	9595	9600	9605	9609	9614	9619	9624	9628	9633	0	1	1	2	2	3	3	4	4
92	9638	9643	9647	9652	9657	9661	9666	9671	9675	9680	0	1	1	2	2	3	3	4	4
93	9685	9689	9694	9699	9703	9708	9713	9717	9722	9727	0	1	1	2	2	3	3	4	4
94	9731	9736	9741	9745	9750	9754	9759	9763	9768	9773	0	1	1	2	2	3	3	4	4
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96	9823	9827	9832	9836	9841	9845	9850	9854	9859	9863	0	1	1	2	2	3	3	4	4
97	9868	9872	9877	9881	9886	9890	9894	9899	9903	9908	0	1	1	2	2	3	3	4	4
98	9912	9917	9921	9926	9930	9934	9939	9943	9948	9952	0	1	1	2	2	3	3	4	4
99	9956	9961	9965	9969	9974	9978	9983	9987	9991	9996	0	1	1	2	2	3	3	3	4
	0	1	2	3	4	5	6	7	8	9	1	2	3	4	5	6	7	8	9

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